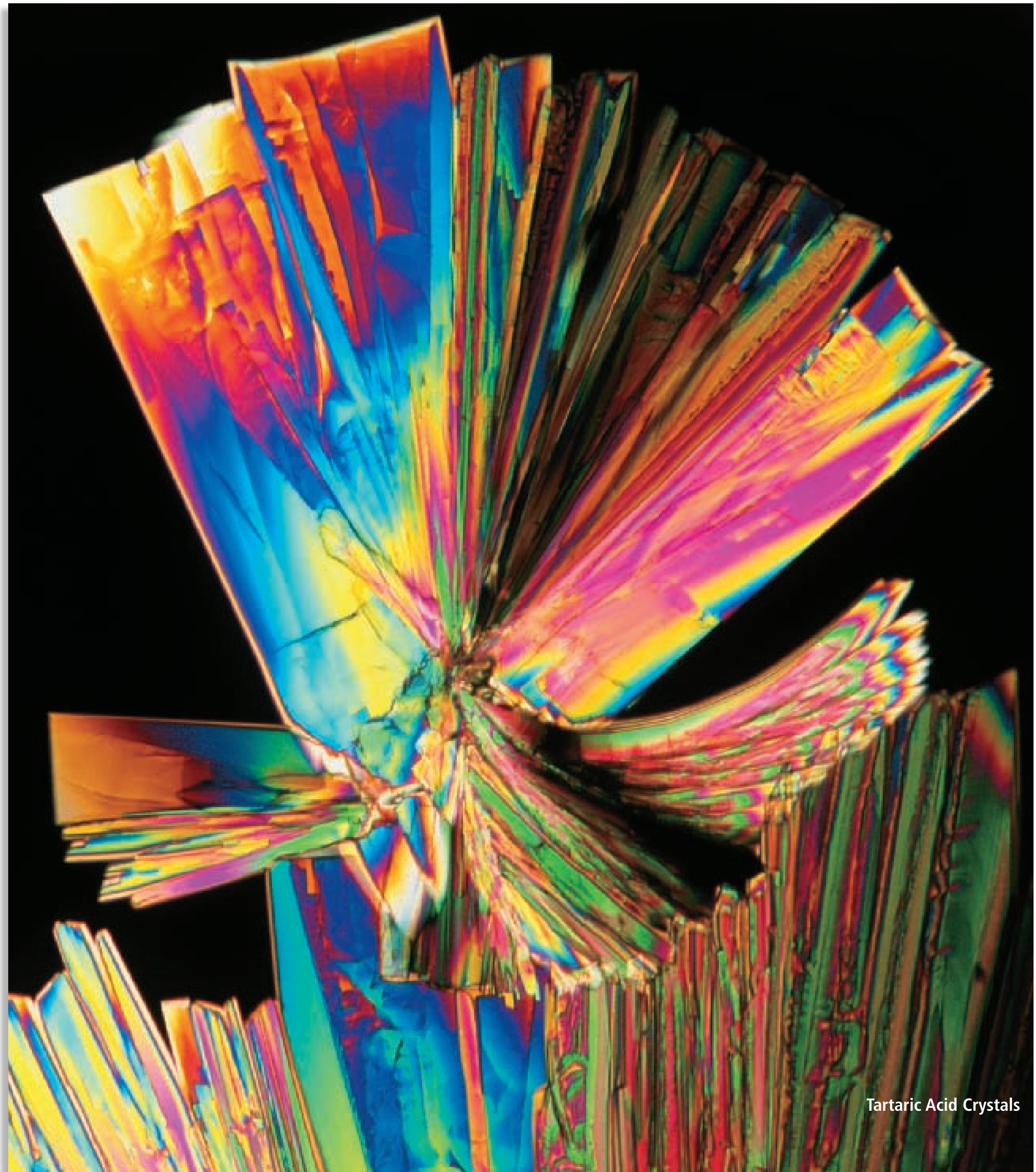
CHAPTER 1

Matter and Change

Chemistry is central to all of the sciences.



Tartaric Acid Crystals

Chemistry Is a Physical Science

SECTION 1

OBJECTIVES

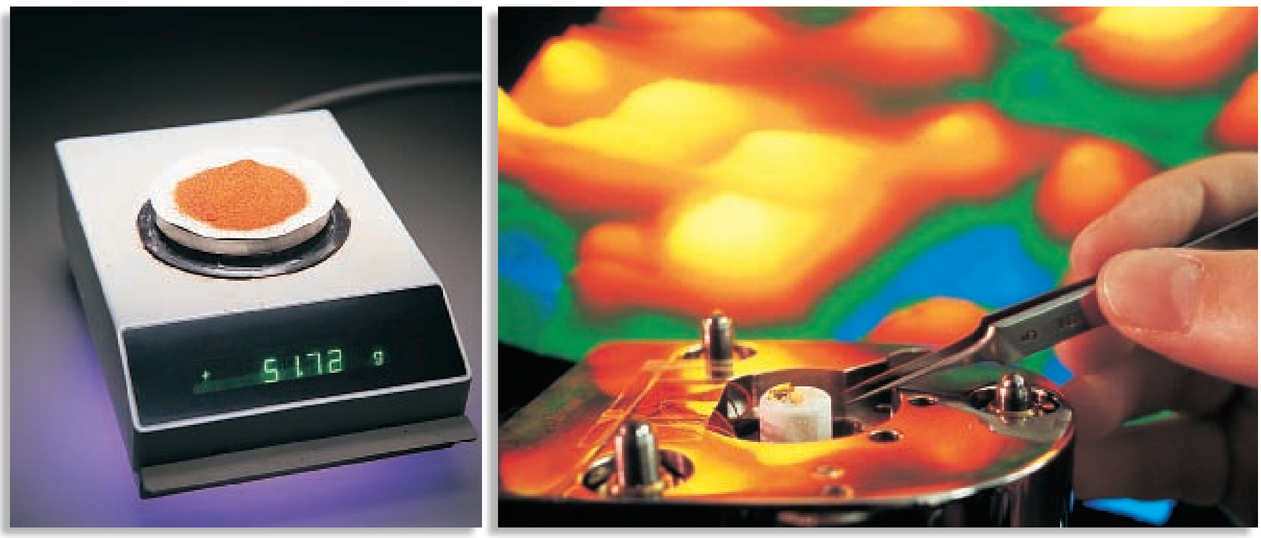
- Define *chemistry.*
- List examples of the branches of chemistry.
- Compare and contrast basic research, applied research, and technological development.

The natural sciences were once divided into two broad categories: the biological sciences and the physical sciences. Living things are the main focus of the biological sciences. The physical sciences focus mainly on nonliving things. However, because we now know that both living and nonliving matter consist of chemical structures, chemistry is central to all the sciences, and there are no longer distinct divisions between the biological and physical sciences.

Chemistry *is the study of the composition, structure, and properties of matter, the processes that matter undergoes, and the energy changes that accompany these processes*. Chemistry deals with questions such as, What is a material's makeup? How does a material change when heated, cooled, or mixed with other materials and why does this behavior occur? Chemists answer these kinds of questions during their work.

Instruments are routinely used in chemistry to extend our ability to observe and make measurements. Instruments make it possible, for example, to look at microstructures—things too tiny to be seen with the unaided eye. The scanning electron microscope reveals tiny structures by beaming particles called electrons at materials. When the electrons hit a material, they scatter and produce a pattern that shows the material's microstructure. Invisible rays called X rays can also be used to

FIGURE 1 A balance (a) is an instrument used to measure the mass of materials. A sample of DNA placed in a scanning tunneling microscope produces an image (b) showing the contours of the DNA's surface.



(a) (b)

determine microstructures. The patterns that appear, called X-ray diffraction patterns, can be analyzed to reveal the arrangement of atoms, molecules, or other particles that make up the material. By learning about microstructures, chemists can explain the behavior of macrostructures—the visible things all around you.

Branches of Chemistry

Chemistry includes many different branches of study and research. The following are six main areas, or branches, of study. But like the biological and physical sciences, these branches often overlap.

1. *Organic chemistry*—the study of most carbon-containing compounds
2. *Inorganic chemistry*—the study of non-organic substances, many of which have organic fragments bonded to metals (organometallics)
3. *Physical chemistry*—the study of the properties and changes of matter and their relation to energy
4. *Analytical chemistry*—the identification of the components and composition of materials
5. *Biochemistry*—the study of substances and processes occurring in living things
6. *Theoretical chemistry*—the use of mathematics and computers to understand the principles behind observed chemical behavior and to design and predict the properties of new compounds

In all areas of chemistry, scientists work with chemicals. *A* **chemical** *is any substance that has a definite composition.* For example, consider the material called sucrose, or cane sugar. It has a definite composition in terms of the atoms that compose it. It is produced by certain plants in the chemical process of photosynthesis. Sucrose is a chemical. Carbon dioxide, water, and countless other substances are chemicals as well.

Knowing the properties of chemicals allows chemists to find suitable uses for them. For example, researchers have synthesized new substances, such as artificial sweeteners and synthetic fibers. The reactions used to make these chemicals can often be carried out on a large scale to make new consumer products such as flavor enhancers and fabrics.

Basic Research

Basic research is carried out for the sake of increasing knowledge, such as how and why a specific reaction occurs and what the properties of a substance are. Chance discoveries can be the result of basic research. The properties of Teflon™, for example, were first discovered by accident. A researcher named Roy Plunkett was puzzled by the fact that a gas cylinder used for an experiment appeared to be empty even though the measured mass of the cylinder clearly indicated there was something inside. Plunkett cut the cylinder open and found a white solid. Through basic research, Plunkett's research team determined the nonstick properties, molecular structure, and chemical composition of the new material.

Applied Research

Applied research is generally carried out to solve a problem. For example, when certain refrigerants escape into the upper atmosphere, they damage the ozone layer, which helps block harmful ultraviolet rays from reaching the surface of Earth. In response to concerns that this atmospheric damage could pose health problems, chemists have developed new refrigerants. In applied research, researchers are driven not by curiosity or a desire to know but by a desire to solve a specific problem.

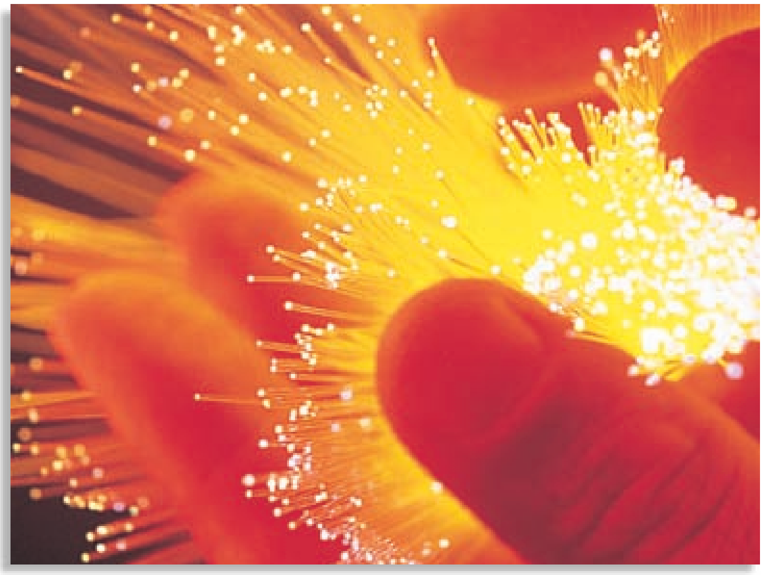


FIGURE 2 The chemical structure of the material in an optical fiber gives it the property of total internal reflection. This property, which allows these fibers to carry light, was discovered through basic and applied research. The use of this property to build networks by sending data on light pulses is the technological development of fiber optics.

Technological Development

Technological development typically involves the production and use of products that improve our quality of life. Examples include computers, catalytic converters for cars, and biodegradable materials.

Technological applications often lag far behind the discoveries that are eventually used in technologies. For example, nonstick cookware, a technological application, was developed well after the accidental discovery of Teflon. When it was later discovered that the Teflon coating on cookware often peeled off, a new challenge arose. Using applied research, scientists were then able to improve the bond between the Teflon and the metal surface of the cookware so that it did not peel.

Basic research, applied research, and technological development often overlap. Discoveries made in basic research may lead to applications that can result in new technologies. For example, knowledge of crystals and light that was gained from basic research was used to develop lasers. It was then discovered that pulses of light from lasers can be sent through optical fibers. Today, telephone messages and cable television signals are carried quickly over long distances using fiber optics.

SECTION REVIEW

1. Define *chemistry*.
2. Name six branches of study in chemistry.
3. Compare and contrast basic research, applied research, and technological development.

Critical Thinking

4. **INFERRING RELATIONSHIPS** Scientific and technological advances are constantly changing how people live and work. Discuss a change that you have observed in your lifetime and that has made life easier or more enjoyable for you.

SECTION 2

Matter and Its Properties

Objectives

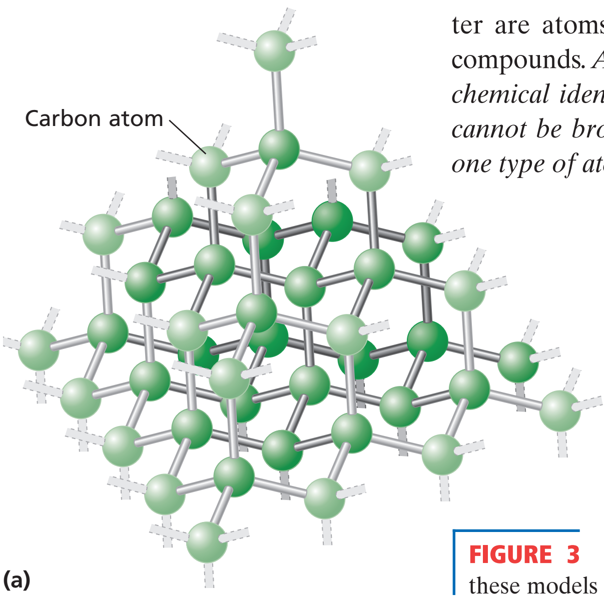
- Distinguish between the physical properties and chemical properties of matter.
- Classify changes of matter as physical or chemical.
- Explain the gas, liquid, and solid states in terms of particles.
- Explain how the law of conservation of energy applies to changes of matter.
- Distinguish between a mixture and a pure substance.

Look around you. You can see a variety of objects—books, desks, chairs, and perhaps trees or buildings outside. All those things are made up of matter, but exactly what is matter? What characteristics, or properties, make matter what it is? In this section, you will learn the answers to these questions.

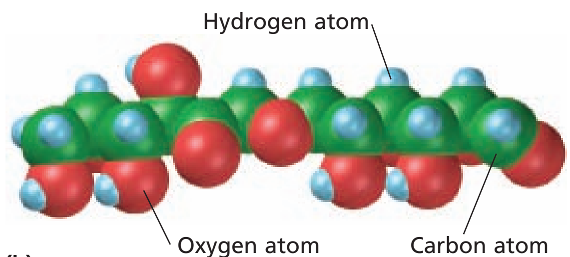
Explaining what matter is involves finding properties that all matter has in common. That may seem difficult, given that matter takes so many different forms. For the moment, just consider one example of matter—a rock. The first thing you might notice is that the rock takes up space. In other words, it has *volume*. Volume is the amount of three-dimensional space an object occupies. All matter has volume. All matter also has a property called mass. **Mass** *is a measure of the amount of matter*. Mass is the measurement you make using a balance. **Matter** can thus be defined as *anything that has mass and takes up space*. These two properties are the general properties of all matter.

Basic Building Blocks of Matter

Matter comes in many forms. The fundamental building blocks of matter are atoms and molecules. These particles make up elements and compounds. *An* **atom** *is the smallest unit of an element that maintains the chemical identity of that element. An* **element** *is a pure substance that cannot be broken down into simpler, stable substances and is made of one type of atom.* Carbon is an element and contains one kind of atom.



(a)



(b)

FIGURE 3 Both elements and compounds are made of atoms, as shown in these models of (a) diamond and (b) sucrose (table sugar).

A **compound** *is a substance that can be broken down into simple stable substances. Each compound is made from the atoms of two or more elements that are chemically bonded.* Water is an example of a compound. It is made of two elements, hydrogen and oxygen. The atoms of hydrogen and oxygen are chemically bonded to form a water molecule. You will learn more about the particles that make up compounds when you study chemical bonding in Chapter 6. For now, you can think of a *molecule* as the smallest unit of an element or compound that retains all of the properties of that element or compound.

Properties and Changes in Matter

Every substance, whether it is an element or a compound, has characteristic properties. Chemists use properties to distinguish between substances and to separate them. Most chemical investigations are related to or depend on the properties of substances.

A property may be a characteristic that defines an entire set of substances. That property can be used to classify an unknown substance as a member of that group. For example, many elements are classified as metals. The distinguishing property of metals is that they conduct electricity well. Therefore, if an unknown element is tested and found to conduct electricity well, it is a metal.

Properties can help reveal the identity of an unknown substance. However, conclusive identification usually cannot be made based on only one property. Comparisons of several properties can be used together to establish the identity of an unknown. Properties are either intensive or extensive. **Extensive properties** *depend on the amount of matter that is present*. Such properties include volume, mass, and the amount of energy in a substance. In contrast, **intensive properties** *do not depend on the amount of matter present*. Such properties include the melting point, boiling point, density, and ability to conduct electricity and to transfer energy as heat. Intensive properties are the same for a given substance regardless of how much of the substance is present. Properties can also be grouped into two general types: physical properties and chemical properties.

Physical Properties and Physical Changes

A **physical property** *is a characteristic that can be observed or measured without changing the identity of the substance*. Physical properties describe the substance itself, rather than describing how it can change into other substances. Examples of physical properties are melting point and boiling point. Those points are, respectively, the temperature at which a substance melts from solid to liquid and the temperature at which it boils from liquid to gas. For example, water melts from ice to liquid at 0°C (273 K or 32°F). Liquid water boils to vapor at 100°C (373 K or 212°F).

A change in a substance that does not involve a change in the identity of the substance is called a **physical change.** Examples of physical

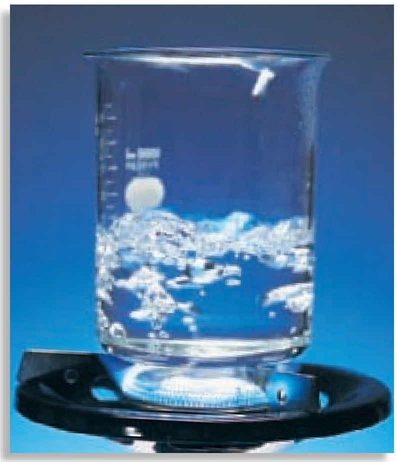


FIGURE 4 Water boils at 100°C no matter how much water is in the container. Boiling point is an intensive property.

changes include grinding, cutting, melting, and boiling a material. These types of changes do not change the identity of the substance present.

Melting and boiling are part of an important class of physical changes called changes of state. As the name suggests, *a* **change of state** *is a physical change of a substance from one state to another*. The three common states of matter are solid, liquid, and gas.

Matter in the **solid** *state has definite volume and definite shape.* For example, a piece of quartz or coal keeps its size and its shape, regardless of the container it is in. Solids have this characteristic because the particles in them are packed together in relatively fixed positions. The particles are held close together by the strong attractive forces between them, and only vibrate about fixed points.

Matter in the **liquid** *state has a definite volume but an indefinite shape;* a liquid assumes the shape of its container. For example, a given quantity of liquid water takes up a definite amount of space, but the water takes the shape of its container. Liquids have this characteristic because the particles in them are close together but can move past one another. The particles in a liquid move more rapidly than those in a solid. This causes them to overcome temporarily the strong attractive forces between them, allowing the liquid to flow.

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Matter in the **gas** *state has neither definite volume nor definite shape.* For example, a given quantity of helium expands to fill any size container and takes the shape of the container. All gases have this characteristic because they are composed of particles that move very rapidly and are at great distances from one another compared with the particles of liquids and solids. At these great distances, the attractive forces between gas particles have less of an effect than they do at the small distances between particles of liquids and solids.

An important fourth state of matter is **plasma.** Plasma is a *high-temperature physical state of matter in which atoms lose most of their electrons, particles that make up atoms.* Plasma is found in fluorescent bulbs.

Melting, the change from solid to liquid, is an example of a change of state. Boiling is a change of state from liquid to gas. Freezing, the opposite of melting, is the change from a liquid to a solid. A change of state does not affect the identity of the substance. For example, when ice melts to liquid water or when liquid water boils to form water vapor, the same substance, water, is still present, as shown in **Figure 6.** The water has simply changed state, but it has not turned into a different compound. Only the distances and interactions between the particles that make up water have changed.

FIGURE 5 Because it possesses certain chemical properties, a test strip containing Benedict's solution is used to test for the presence of sugar in urine. The test strip is dipped into the sample. The test strip is then matched to a color scale to determine the sugar level in the urine.

Chemical Properties and Chemical Changes

Physical properties can be observed without changing the identity of the substance, but properties of the second type—chemical properties—cannot. *A* **chemical property** *relates to a substance's ability to undergo changes that transform it into different substances.* Chemical properties are easiest to see when

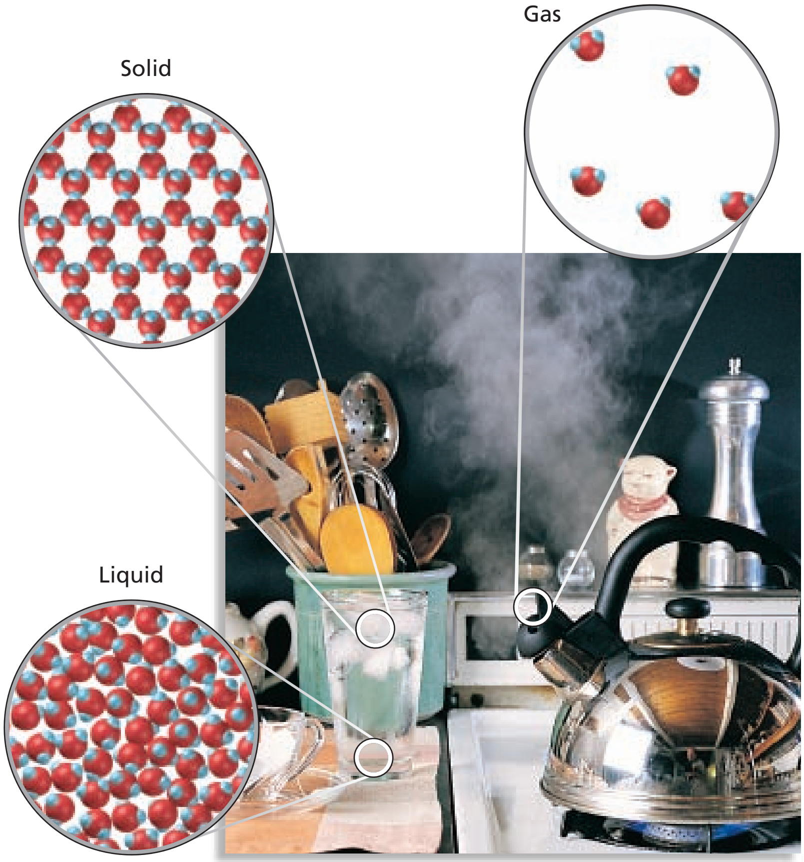


FIGURE 6 Models for water in three states. The molecules are close together in the solid and liquid states but far apart in the gas state. The molecules in the solid state are relatively fixed in position, but those in the liquid and gas states can flow around each other.

substances react to form new substances. For example, the ability of charcoal (carbon) to burn in air is a chemical property. When charcoal burns, it combines with oxygen in air to become a new substance, carbon dioxide gas. After the chemical change, the amounts of the original substances, carbon and oxygen, are less than before. A different substance with different properties has been formed. Other examples of chemical properties include the ability of iron to rust by combining with oxygen in air and the ability of silver to tarnish by combining with sulfur.

A change in which one or more substances are converted into different substances is called a **chemical change** *or* **chemical reaction.** *The substances that react in a chemical change are called the* **reactants.** *The substances that are formed by the chemical change are called the* **products.** In the case of burning charcoal, carbon and oxygen are the reactants in a combustion, or burning, reaction. Carbon dioxide is the product. The chemical change can be described as follows:

Carbon plus oxygen yields (or forms) carbon dioxide.

Arrows and plus signs can be substituted for the words *yields* and *plus,* respectively:

$$\text{carbon} + \text{oxygen} \longrightarrow \text{carbon dioxide}$$

extension

Historical Chemistry

Go to **go.hrw.com** for a full-length article on the chemical reactions of noble gases.

Keyword: HC6MTXX

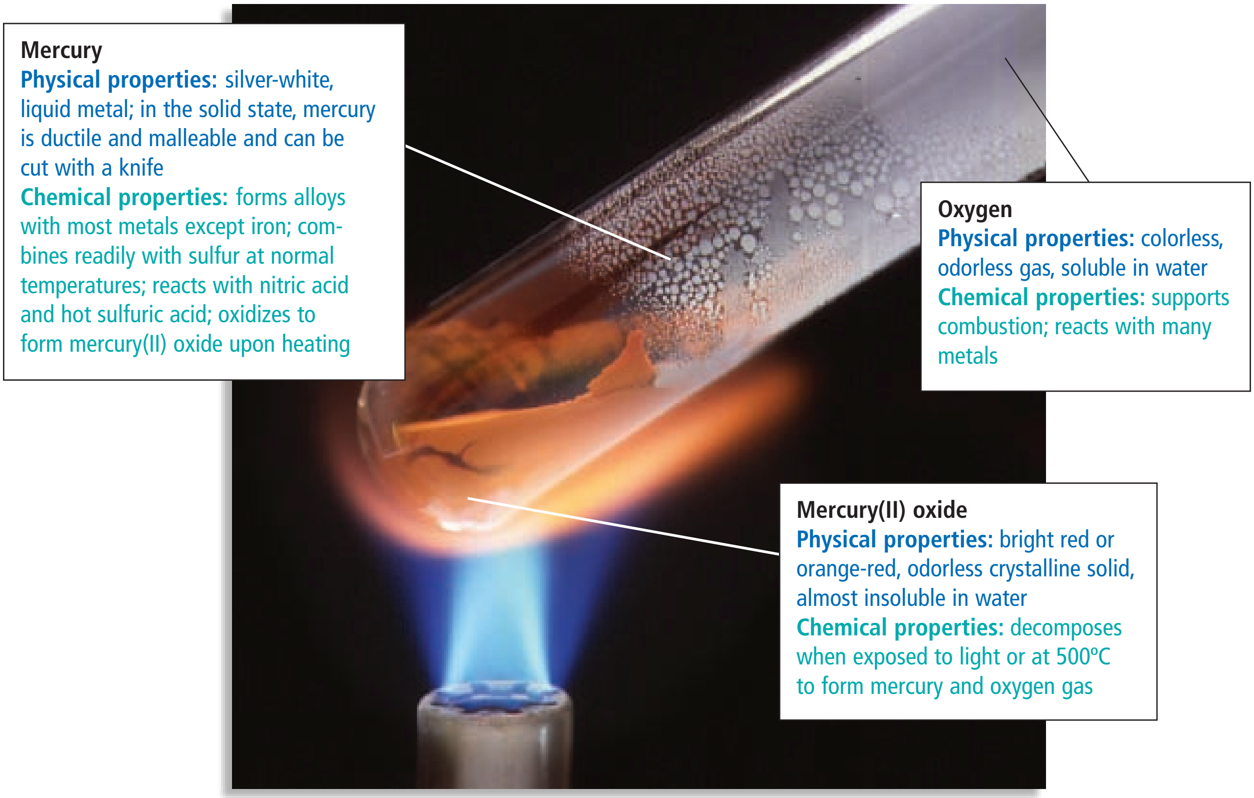


FIGURE 7 When mercury(II) oxide is heated, it decomposes to form oxygen gas and mercury (which can be seen on the side of the test tube). Decomposition is a chemical change that can be observed by comparing the properties of mercury(II) oxide, mercury, and oxygen.

The decomposition of the mercury compound shown in **Figure 7** can be expressed as follows:

$$\text{mercury(II) oxide} \longrightarrow \text{mercury} + \text{oxygen}$$

Chemical changes and reactions, such as combustion and decomposition, form products whose properties differ greatly from those of the reactants. However, chemical changes do not affect the total amount of matter present before and after a reaction. The amount of matter, and therefore the total mass, remains the same.

Energy and Changes in Matter

When physical or chemical changes occur, energy is always involved. The energy can take several different forms, such as heat or light. Sometimes heat provides enough energy to cause a physical change, as in the melting of ice, and sometimes heat provides enough energy to cause a chemical change, as in the decomposition of water vapor to form oxygen gas and hydrogen gas. But the boundary between physical and chemical changes isn't always so clear. For example, while most chemists would consider the dissolving of sucrose in water to be a physical change, many chemists would consider the dissolving of table salt in water to be a chemical change. As you learn more about the structure of matter, you will better understand why the boundaries between chemical and physical changes can be confusing.

Accounting for all the energy present before and after a change is not a simple process. But scientists who have done such experimentation are confident that the total amount of energy remains the same. Although energy can be absorbed or released in a change, it is not destroyed or created. It simply assumes a different form. This is the law of conservation of energy.

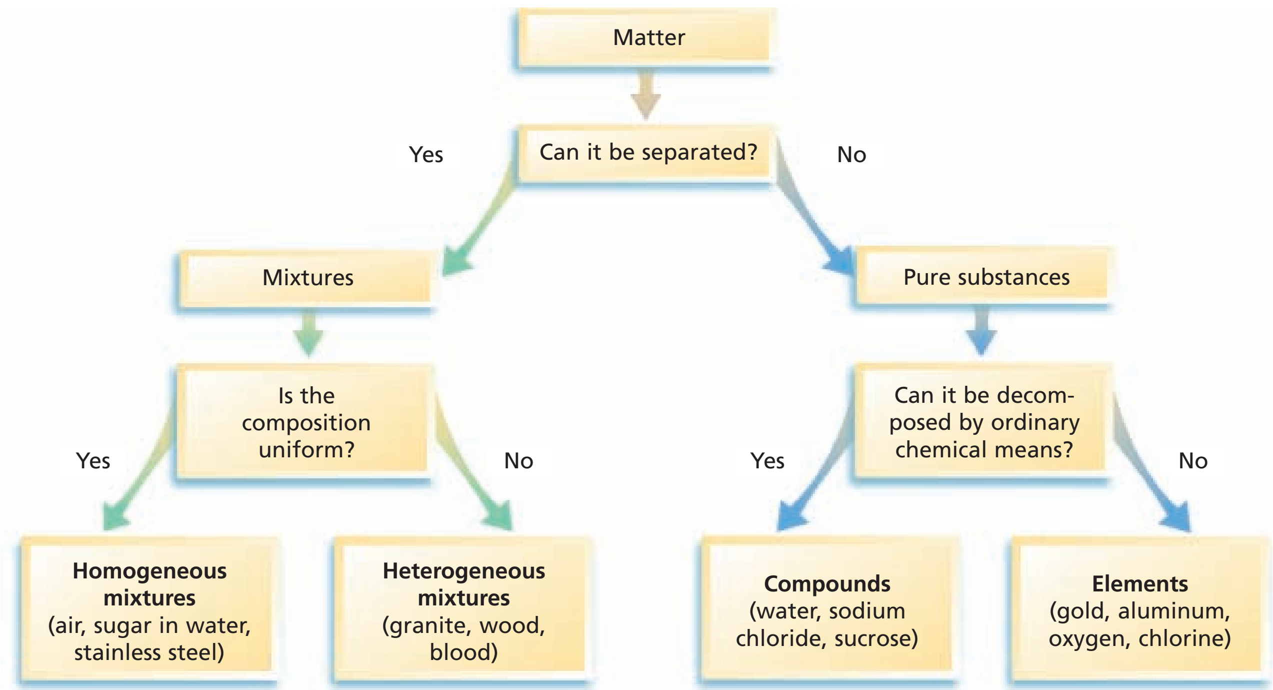
Classification of Matter

Matter exists in an enormous variety of forms. Any sample of matter, however, can be classified either as a pure substance or as a mixture. The composition of a pure substance is the same throughout and does not vary from sample to sample. A pure substance can be an element or a compound. Mixtures, in contrast, contain more than one substance. They can vary in composition and properties from sample to sample and sometimes from one part of a sample to another part of the same sample. All matter, whether it is a pure substance or a mixture, can be classified in terms of uniformity of composition and properties of a given sample. **Figure 8** illustrates the overall classification of matter into elements, compounds, and mixtures.

Mixtures

You deal with mixtures every day. Nearly every object around you, including most things you eat and drink and even the air you breathe, is a mixture. *A* **mixture** *is a blend of two or more kinds of matter, each*

FIGURE 8 This classification scheme for matter shows the relationships among mixtures, compounds, and elements.



(a) (b) (c)

FIGURE 9 (a) Barium chromate can be separated from the solution in the beaker using filtration. (b) A centrifuge can be used to separate certain solid components. The centrifuge spins rapidly, which causes the solids to settle to the bottom of the test tube. (c) The components of an ink can be separated using paper chromatography.

of which retains its own identity and properties. The parts, or components, of a mixture are simply mixed together physically and can usually be separated. As a result, the properties of a mixture are a combination of the properties of its components. Because mixtures can contain various amounts of different substances, a mixture's composition must be specified. This is often done in terms of percentage by mass or by volume. For example, a mixture might be 5% sodium chloride and 95% water by mass.

Some mixtures are *uniform in composition;* that is, they are said to be **homogeneous.** They have the same proportion of components throughout. *Homogeneous mixtures are also called* **solutions.** A salt-water solution is an example of such a mixture. Other mixtures are *not uniform throughout;* that is, they are **heterogeneous.** For example, in a mixture of clay and water, heavier clay particles concentrate near the bottom of the container.

Some mixtures can be separated by filtration or vaporized to separate the different components. Filtration can be used to separate a mixture of solid barium chromate from the other substances, as shown in the beaker in **Figure 9a.** The yellow barium compound is trapped by the filter paper, but the solution passes through. If the solid in a liquid-solid mixture settles to the bottom of the container, the liquid can be carefully poured off (decanted). A centrifuge **(Figure 9b)** can be used to separate some solid-liquid mixtures, such as those in blood. Another technique, called paper chromatography, can be used to separate mixtures of dyes or pigments because the different substances move at different rates on the paper **(Figure 9c).**

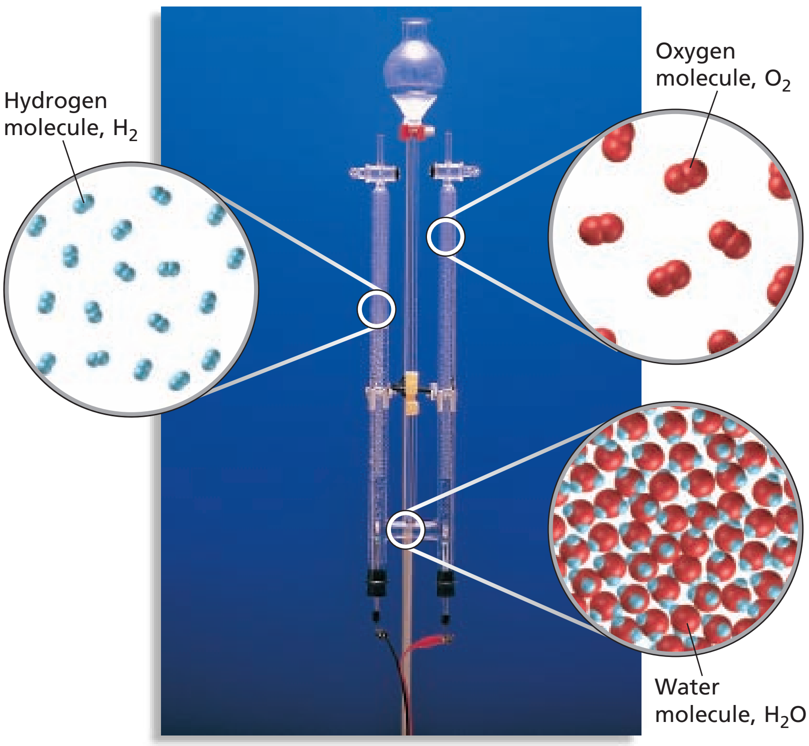
Pure Substances

Any sample of a pure substance is homogeneous. *A* **pure substance** *has a fixed composition and differs from a mixture in the following ways:*

1. *Every sample of a given pure substance has exactly the same characteristic properties.* All samples of a pure substance have the same characteristic physical and chemical properties. These properties are so specific that they can be used to identify the substance. In contrast, the properties of a mixture depend on the relative amounts of the mixture's components.
2. *Every sample of a given pure substance has exactly the same composition.* Unlike mixtures, all samples of a pure substance have the same makeup. For example, pure water is always 11.2% hydrogen and 88.8% oxygen by mass.

Pure substances are either compounds or elements. A compound can be decomposed, or broken down, into two or more simpler compounds or elements by a chemical change. Water is a compound made of hydrogen and oxygen chemically bonded to form a single substance. Water can be broken down into hydrogen and oxygen through a chemical reaction called electrolysis, as shown in **Figure 10a.**

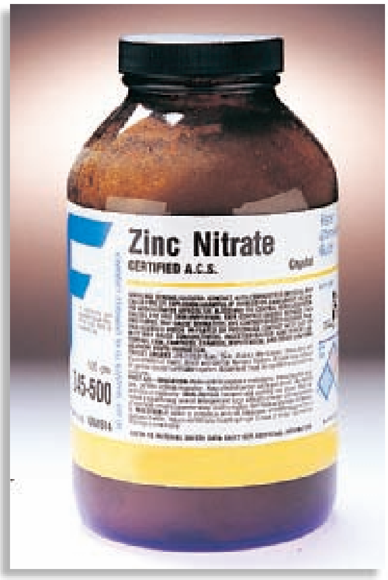
Sucrose is made of carbon, hydrogen, and oxygen. Sucrose breaks down to form the other substances shown in **Figure 10b.** Under intense heating, sucrose breaks down to produce carbon and water.



(a)

(b)

FIGURE 10 (a) Passing an electric current through water causes the compound to break down into the elements hydrogen and oxygen, which differ in composition from water. (b) When sucrose is heated, it caramelizes. When it is heated to a high enough temperature, it breaks down completely into carbon and water.



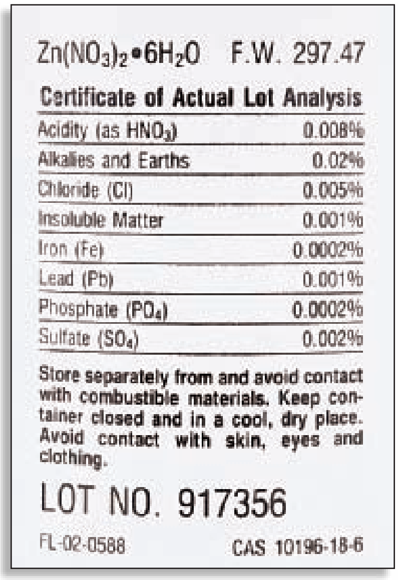


FIGURE 11 The labeling on a reagent bottle lists the grade of the reagent and the percentages of impurities for that grade. What grade is this chemical?

TABLE 1 ***Some Grades of Chemical Purity***

Increasing purity (↑)
Primary standard reagents
ACS (American Chemical Society–specified reagents)
USP (United States Pharmacopoeia standards)
CP (chemically pure; purer than technical grade)
NF (National Formulary specifications)
FCC (Food Chemical Code specifications)
Technical (industrial chemicals)

Laboratory Chemicals and Purity

The chemicals in laboratories are generally treated as if they are pure. However, all chemicals have some impurities. Chemical grades of purity are listed in **Table 1.** The purity ranking of the grades can vary when agencies differ in their standards. For some chemicals, the USP grade may specify higher purity than the CP grade. For other chemicals, the opposite may be true. However, the primary standard reagent grade is always purer than the technical grade for the same chemical. Chemists need to be aware of the kinds of impurities in a reagent because these impurities could affect the results of a reaction. For example, the chemical label shown in **Figure 11** shows the impurities for that grade. The chemical manufacturer must ensure that the standards set for that reagent by the American Chemical Society are met.

SECTION REVIEW

1. **a.** What is the main difference between physical properties and chemical properties?
b. Give an example of each.

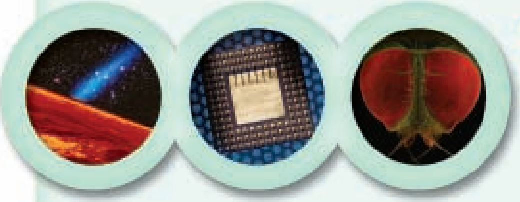
2. Classify each of the following as either a physical change or a chemical change.
a. tearing a sheet of paper
b. melting a piece of wax
c. burning a log

3. How do you decide whether a sample of matter is a solid, liquid, or gas?

4. Contrast mixtures with pure substances.

Critical Thinking

5. **ANALYZING INFORMATION** Compare the composition of sucrose purified from sugar cane with the composition of sucrose purified from sugar beets. Explain your answer.



CROSS-DISCIPLINARY CONNECTION

Secrets of the Cremona Violins

What are the most beautiful sounding of all violins? Most professionals will pick the instruments created in Cremona, Italy, following the Renaissance. At that time, Antonio Stradivari, the Guarneri family, and other designers created instruments of extraordinary sound that have yet to be matched. The craftsmen were notoriously secretive about their techniques, but, based on 20 years of research, Dr. Joseph Nagyvary, a professor of biochemistry at Texas A&M University, thinks he has discovered the key to the violins' sound hidden in the chemistry of their materials.

According to Dr. Nagyvary, Stradivarius instruments are nearly free of the shrill, high-pitched noises produced by modern violins. Generally, violin makers attribute this to the design of the instrument, but Dr. Nagyvary traces it to a different source. In Stradivari's day, wood for the violins was transported by floating it down a river from the mountains to Venice, where it was stored in sea water. Dr. Nagyvary first theorized that the soaking process could have removed ingredients from the wood that made it inherently noisy. His experiments revealed that microbes and minerals also permeated the wood, making their own contribution to the mellow musical sound. Attempting to reproduce the effects of sea water, Dr. Nagyvary soaks all his wood in a "secret" solution. One of his favorite ingredients is a cherry-and-plum puree,

▲*Dr. Nagyvary and his violin*

which contains an enzyme called pectinase. The pectinase softens the wood, making it resonate more freely.

"The other key factor in a violin's sound," says Dr. Nagyvary, "is the finish, which is the filler and the varnish covering the instrument. Most modern finishes are made from rubbery materials, which limit the vibrations of the wood." Modern analysis has revealed that the Cremona finish was different: it was a brittle mineral microcomposite of a very sophisticated nature. According to historical accounts, all violin makers, including Stradivari, procured their varnishes from the local drugstore chemist, and they didn't even know what they were using! Dr. Nagyvary and his co-workers have identified most of the key ingredients of the Cremona finish.

Many new violins made from the treated wood and replicated finish have been made, and their sound has been analyzed by modern signal analyzers. These violins have been favorably compared with authentic Stradivari violins.

A number of expert violinists have praised the sound of Dr. Nagyvary's instruments, but some violin makers remain skeptical of the chemist's claims. They insist that it takes many years to reveal just how good a violin is. In the meantime, almost everyone agrees that the art and science of violin making are still epitomized by the instruments of Cremona.

Questions

1. According to Dr. Nagyvary, what are two factors that are believed to have created the unique sound of the Stradivari violins?

2. Use the library or Internet resources to find additional information about the Cremona violin makers. Who were some of the other instrument makers during the time period in which Stradivari was alive? Were other stringed instruments made by these artisans? What are the estimated present-day values of instruments made during this period in Cremona?

SECTION 3

Elements

Objectives

- Use a periodic table to name elements, given their symbols.
- Use a periodic table to write the symbols of elements, given their names.
- Describe the arrangement of the periodic table.
- List the characteristics that distinguish metals, nonmetals, and metalloids.

As you have read, elements are pure substances that cannot be decomposed by chemical changes. The elements serve as the building blocks of matter. Each element has characteristic properties. The elements are organized into groups based on similar chemical properties. This organization of elements is the *periodic table,* which is shown in **Figure 12** on the next page.

Introduction to the Periodic Table

Each small square of the periodic table shows the symbol for the element and the atomic number. For example, the first square, at the upper left, represents element 1, hydrogen, which has the symbol H. As you look through the table, you will see many familiar elements, including iron, sodium, neon, silver, copper, aluminum, sulfur, and lead. You can often relate the symbols to the English names of the elements. Some symbols are derived from the element's older name, which was often in Latin. Still others come from German. For example, wolfram comes from the German name for tungsten. **Table 2** lists some elements and their older names.

TABLE 2 ***Elements with Symbols Based on Older Names***

Modern name	Symbol	Older name
Antimony	Sb	stibium
Copper	Cu	cuprum
Gold	Au	aurum
Iron	Fe	ferrum
Lead	Pb	plumbum
Mercury	Hg	hydrargyrum
Potassium	K	kalium
Silver	Ag	argentum
Sodium	Na	natrium
Tin	Sn	stannum
Tungsten	W	wolfram

Periodic Table

Metals

Metalloids

Nonmetals

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg							
				58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
				90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

FIGURE 12 The periodic table of elements. The names of the elements can be found on Table A-6 in the appendix.

The vertical columns of the periodic table are called **groups,** or **families.** Notice that they are numbered from 1 to 18 from left to right. Each group contains elements with similar chemical properties. For example, the elements in Group 2 are beryllium, magnesium, calcium, strontium, barium, and radium. All of these elements are reactive metals with similar abilities to bond to other kinds of atoms. The two major categories of elements are metals and nonmetals. Metalloids have properties intermediate between those of metals and nonmetals.

The horizontal rows of elements in the periodic table are called **periods.** Physical and chemical properties change somewhat regularly across a period. Elements that are close to each other in the same period tend to be more similar than elements that are farther apart. For example, in Period 2, the elements lithium and beryllium, in Groups 1 and 2, respectively, are somewhat similar in properties. However, their properties are very different from the properties of fluorine, the Period-2 element in Group 17.

The two sets of elements placed below the periodic table make up what are called the lanthanide series and the actinide series. These metallic elements fit into the table just after elements 57 and 89. They are placed below the table to keep the table from being too wide.

There is a section in the back of this book called the *Elements Handbook* which covers some elements in greater detail. You will use information from the handbook to complete the questions in the Using the Handbook sections in the chapter reviews.

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Topic: Element Names
Code: HC60495

Chemistry in Action

Superconductors

Any metal becomes a better conductor of electrical energy as its temperature decreases. In 1911, scientists discovered that when mercury is cooled to about –269°C, it loses all resistance and becomes a superconductor. Scientists have long tried to find a material that would superconduct at temperatures above –196°C, the boiling point of liquid nitrogen. In 1987, scientists discovered ceramic materials that became superconductors when cooled only to –183°C. These "high-temperature" superconductors are used to build very powerful electromagnets. Ceramic electromagnets are used in medical magnetic resonance imaging (MRI) machines and in high-efficiency electric motors and generators.

Types of Elements

The periodic table is broadly divided into two main sections: metals and nonmetals. As you can see in **Figure 12,** the metals are at the left and in the center of the table. The nonmetals are toward the right. Some elements, such as boron and silicon, show characteristics of both metals and nonmetals.

Metals

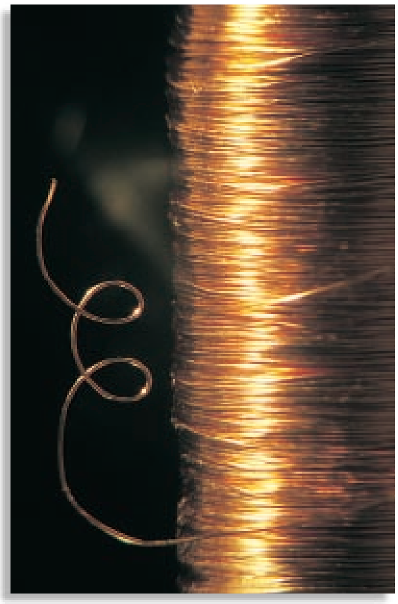
Some of the properties of metals may be familiar to you. For example, you can recognize metals by their shininess, or metallic luster. Perhaps the most important characteristic property of metals is the ease with which they conduct electricity and transfer energy. Thus, *a* **metal** *is an element that is a good electrical conductor and a good heat conductor.*

At room temperature, most metals are solids. Most metals also have the property of *malleability,* that is, they can be hammered or rolled into thin sheets. Metals also tend to be *ductile,* which means that they can be drawn into a fine wire. Metals behave this way because they have high *tensile strength,* the ability to resist breaking when pulled.

Although all metals conduct electricity well, metals also have very diverse properties. Mercury is a liquid at room temperature, whereas tungsten has the highest melting point of any element. The metals in Group 1 are so soft that they can be cut with a knife, yet others, such as chromium, are very hard. Some metals, such as manganese and bismuth, are very brittle, yet others, such as iron and copper, are very malleable and ductile. Most metals have a silvery or grayish white *luster*. Two exceptions are gold and copper, which are yellow and reddish brown, respectively. **Figure 13** shows examples of metals.

FIGURE 13 (a) Gold has a low reactivity, which is why it may be found in nature in relatively pure form. (b) Copper is used in wiring because it is ductile and conducts electrical energy (c) Aluminum is malleable. It can be rolled into foil that is used for wrapping food.

(a)



(b)

(c)

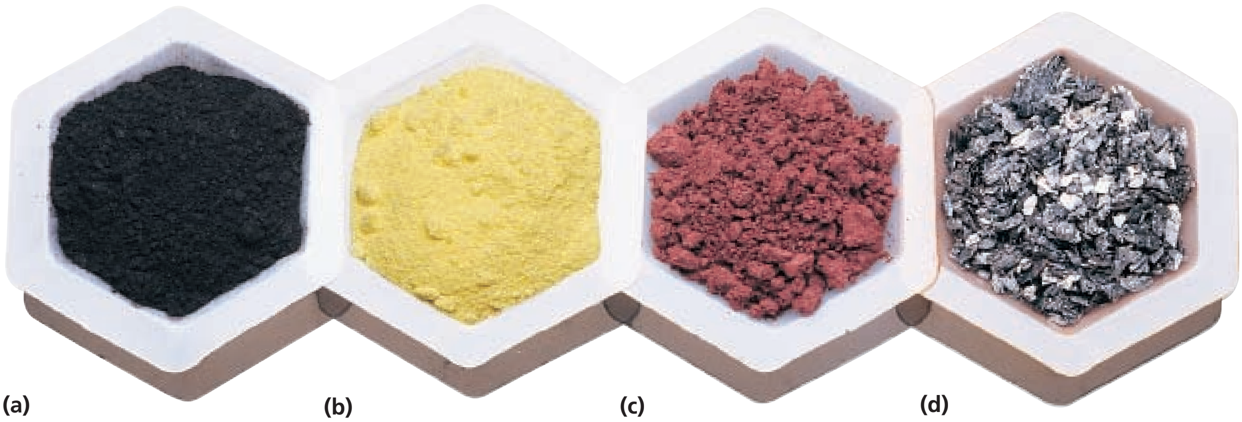


FIGURE 14 Various nonmetallic elements: (a) carbon, (b) sulfur, (c) phosphorus, and (d) iodine

Copper: A Typical Metal

Copper has a characteristic reddish color and a metallic luster. It is found naturally in minerals such as chalcopyrite and malachite. Pure copper melts at 1083°C and boils at 2567°C. It can be readily drawn into fine wire, pressed into thin sheets, and formed into tubing. Copper conducts electricity with little loss of energy.

Copper remains unchanged in pure, dry air at room temperature. When heated, it reacts with oxygen in air. It also reacts with sulfur and the elements in Group 17 of the periodic table. The green coating on a piece of weathered copper comes from the reaction of copper with oxygen, carbon dioxide, and sulfur compounds. Copper is an essential mineral in the human diet.

Nonmetals

Many nonmetals are gases at room temperature. These include nitrogen, oxygen, fluorine, and chlorine. One nonmetal, bromine, is a liquid. The solid nonmetals include carbon, phosphorus, selenium, sulfur, and iodine. These solids tend to be brittle rather than malleable and ductile. Some nonmetals are illustrated in **Figure 14.**

Low conductivity can be used to define nonmetals. *A* **nonmetal** *is an element that is a poor conductor of heat and electricity.* If you look at **Figure 12,** you will see that there are fewer nonmetals than metals.

Phosphorus: A Typical Nonmetal

Phosphorus is one of five solid nonmetals. Pure phosphorus is known in two common forms. Red phosphorus is a dark red powder that melts at 597°C. White phosphorus is a waxy solid that melts at 44°C. Because it ignites in air at room temperature, white phosphorus is stored under water.

Phosphorus is too reactive to exist in pure form in nature. It is present in huge quantities in phosphate rock, where it is combined with oxygen and calcium. All living things contain phosphorus.

Metalloids

As you look from left to right on the periodic table, you can see that the metalloids are found between the metals and the nonmetals. *A* **metalloid**

FIGURE 15 Selenium is a nonmetal, though it looks metallic.

FIGURE 16 Some noble gases are used to make lighted signs of various colors.

is an element that has some characteristics of metals and some characteristics of nonmetals. All metalloids are solids at room temperature. They tend to be less malleable than metals but not as brittle as nonmetals. Some metalloids, such as antimony, have a somewhat metallic luster.

Metalloids tend to be semiconductors of electricity. That is, their ability to conduct electricity is intermediate between that of metals and that of nonmetals. Metalloids are used in the solid state circuitry found in desktop computers, digital watches, televisions, and radios.

Noble Gases

The elements in Group 18 of the periodic table are the noble gases. These elements are generally unreactive. In fact, it was not until 1962 that the first noble gas compound, xenon hexafluoroplatinate, was prepared. Low reactivity makes the noble gases very different from the other families of elements. Group 18 elements are gases at room temperature. Neon, argon, krypton, and xenon are all used in lighting. Helium is used in party balloons and weather balloons because it is less dense than air.

SECTION REVIEW

1. Use the periodic table to write the names for the following elements: O, S, Cu, Ag.
2. Use the periodic table to write the symbols for the following elements: iron, nitrogen, calcium, mercury.
3. Which elements are most likely to undergo the same kinds of reactions, those in a group or those in a period?
4. Describe the main differences between metals, nonmetals, and metalloids.

Critical Thinking

5. **INFERRING CONCLUSIONS** If you find an element in nature in its pure elemental state, what can you infer about the element's chemical reactivity? How can you tell whether that element is a metal or a nonmetal?

CHAPTER HIGHLIGHTS

Chemistry Is a Physical Science

Vocabulary
chemistry
chemical

- Chemistry is the study of the composition, structure, and properties of matter and the changes that matter undergoes.
- A chemical is any substance that has a definite composition or is used or produced in a chemical process.
- Basic research is carried out for the sake of increasing knowledge. Applied research is carried out to solve practical problems. Technological development involves the use of existing knowledge to make life easier or more convenient.

Matter and Its Properties

Vocabulary

mass	change of state
matter	solid
atom	liquid
element	gas
compound	plasma
extensive property	chemical property
intensive property	chemical change
physical property	chemical reaction
physical change	reactant

- All matter has mass and takes up space. Mass is one measure of the amount of matter.
- Chemical properties refer to a substance's ability to undergo changes that alter its composition and identity.
- An element is composed of one kind of atom. Compounds are made from two or more elements in fixed proportions.
- All substances have characteristic properties that enable chemists to tell the substances apart and to separate the substances.
- Physical changes do not involve changes in identity of a substance.
- The three major states of matter are solid, liquid, and gas. Changes of state, such as melting and boiling, are physical changes.
- In a chemical change—or a chemical reaction—the identity of the substance changes.
- Energy changes accompany physical and chemical changes. Energy may be released or absorbed, but it is neither created nor destroyed.
- Matter can be classified into mixtures and pure substances.

Elements

Vocabulary
group
family
period
metal
nonmetal
metalloid

- Each element has a unique symbol. The periodic table shows the elements organized by their chemical properties. Columns on the table represent groups or families of elements that have similar chemical properties. Properties vary across the rows, or periods.
- The elements can be classified as metals, nonmetals, metalloids, and noble gases. These classes occupy different areas of the periodic table. Metals tend to be shiny, malleable, and ductile and tend to be good conductors. Nonmetals tend to be brittle and tend to be poor conductors.
- Metalloids are intermediate in properties between metals and nonmetals, and they tend to be semiconductors of electricity. The noble gases are generally unreactive elements.

CHAPTER REVIEW

Chemistry Is a Physical Science

SECTION 1 REVIEW

1. What is chemistry?

2. What branch of chemistry is most concerned with the study of carbon compounds?

3. What is meant by the word *chemical*, as used by scientists?

4. In which of the six branches of chemistry would a scientist be working if he or she were doing the following:

a. investigating energy relationships for various reactions

b. comparing properties of alcohols with those of sugars

c. studying reactions that occur during the digestion of food

5. Identify each of the following as an example of either basic research, applied research, or technological development:

a. A new type of refrigerant that is less damaging to the environment is developed.

b. A new element is synthesized in a particle accelerator.

c. A computer chip is redesigned to increase the speed of the computer.

Matter and Its Properties

SECTION 2 REVIEW

6. a. What is mass?

b. What is volume?

7. How does the composition of a pure compound differ from that of a mixture?

8. a. Define property.

b. How are properties useful in classifying materials?

9. What is the difference between extensive properties and intensive properties?

10. a. Define chemical property.

b. List two examples of chemical properties.

11. Distinguish between a physical change and a chemical change.

12. a. How does a solid differ from a liquid?

b. How does a liquid differ from a gas?

c. How is a liquid similar to a gas?

d. What is a plasma?

13. What is meant by a change in state?

14. Identify the reactants and products in the following reaction:

potassium + water $\longrightarrow$
potassium hydroxide + hydrogen

15. Suppose different parts of a sample material have different compositions. What can you conclude about the material?

Elements

SECTION 3 REVIEW

16. What is the significance of the vertical columns of the periodic table? What is the significance of the horizontal rows?

17. Compare the physical properties of metals, nonmetals, metalloids, and noble gases, and describe where in the periodic table each of these kinds of elements is located.

18. Suppose element X is a poor conductor of electricity and breaks when hit with a hammer. Element Z is a good conductor of electricity and heat. In what area of the periodic table does each element most likely belong?

19. Use the periodic table to write the names of the elements that have the following symbols, and identify each as a metal, nonmetal, metalloid, or noble gas.

a. K	c. Si	e. Hg
b. Ag	d. Na	f. He

20. An unknown element is shiny and is found to be a good conductor of electricity. What other properties would you predict for it?

21. Use the periodic table to identify the group numbers and period numbers of the following elements:

a. carbon, C	c. chromium, Cr
b. argon, Ar	d. barium, Ba

MIXED REVIEW

22. a. Define physical property.
b. List two examples of physical properties.

23. How can you tell the difference between an element and a compound?

24. Identify each of the following as either a physical change or a chemical change. Explain your answers.
a. A piece of wood is sawed in half.
b. Milk turns sour.
c. Melted butter solidifies in the refrigerator.

25. Write a brief paragraph that shows that you understand the following terms and the relationships between them: *atom, molecule, compound,* and *element.*

26. Pick an object you can see right now. List three of the object's physical properties that you can observe. Can you also observe a chemical property of the object? Explain your answer.

CRITICAL THINKING

27. **Interpreting Concepts** One way to make lemonade is to start by combining lemon juice and water. To make the lemonade taste better you could add some sugar. Is your lemonade-sugar combination classified as a compound or a mixture? Explain your answer.

28. **Analyzing Results** A pure white, solid material that looks like table salt releases gas when heated under certain conditions. There is no change in the appearance of the solid, but the reactivity of the material changes.
a. Did a chemical or physical change occur? How do you know?
b. Was the original material an element or a compound?

29. **Interpreting Concepts**
a. Is breaking an egg an example of a physical or chemical change? Explain your answer.
b. Is cooking an egg an example of a physical or chemical change? Explain your answer.

USING THE HANDBOOK

30. Review the information on trace elements in the *Elements Handbook* in the back of this text.
a. What are the functions of trace elements in the body?
b. What transition metal plays an important role in oxygen transport throughout the body?
c. What two Group 1 elements are part of the electrolyte balance in the body?

RESEARCH & WRITING

31. Research any current technological product of your choosing. Find out about its manufacture and uses. Also find out about the basic research and applied research that made its development possible.

32. Investigate current and proposed technological applications of superconductors. Find out which of these applications have been successfully tested or are already in use.

ALTERNATIVE ASSESSMENT

33. During a 1 h period, make a list of all the changes that you see around you and that involve matter. Note whether each change seems to be a physical change or a chemical change. Give reasons for your answers.

34. Make a concept map using at least 15 terms from the vocabulary lists. An introduction to concept mapping is found in **Appendix B** of this book.

extension

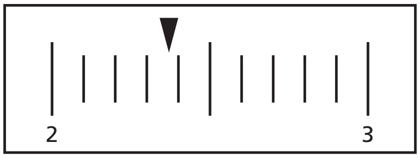
Graphing Calculator
Graphing Tabular Data
Go to **go.hrw.com** for a graphing calculator exercise that asks you to graph temperature vs. time for a chemical reaction.

Keyword : HC6MTXX

Math Tutor SIGNIFICANT FIGURES

The certainty of a measurement is expressed by significant figures. Significant figures in a measurement consist of all the digits known with certainty plus one final digit. Look at the reading below, which was obtained when measuring the mass of a paper clip.

mass of paperclip



Balance reading = 2.37 g

You know with certainty that the paper clip has a mass of 2.3 g. You can also estimate an additional mass of 0.07 g for a total of 2.37 g. Each of the three digits in 2.37 g is significant because it is either certain or estimated.

Problem-Solving TIPS

- Every nonzero digit is significant. Zeros between nonzero digits are significant.
- Zeros appearing in front of the first nonzero digit are not significant.
- If there is no decimal point, zeros that follow the last nonzero digit are not significant.
- If there is a decimal point, zeros that follow the last nonzero digit are significant.
- When measurements are added or subtracted, the result must be rounded to the same number of decimal places that the quantity with the fewest decimal places has.
- When measurements are multiplied or divided, the result must be rounded to the same number of significant figures that the quantity with the smallest number of significant figures has.

SAMPLE 1

How many significant figures does 0.007 09 kg have?

All nonzero digits are significant. The zero between the 7 and 9 is significant. The zeros to the left of the decimal point are not significant. The quantity 0.007 09 kg has 3 significant figures.

SAMPLE 2

Divide 79.7 g by 0.89 cm^3.

The quantity 79.7 g has 3 significant figures, but 0.89 cm^3 has only 2 significant figures. So, the product 8.955 056 18 g/cm^3 must be rounded to 2 significant figures. The rounded quantity is 9.0 g/cm^3.

PRACTICE PROBLEMS

1. Determine the number of significant figures.

a. 42.200 L
b. 0.055 00 mol

2. Perform the following calculations and apply the rules for significant figures.

a. 56.05 g ÷ 13.3 cm^3 b. 1.057 g + 3.02 g + 12.4 g

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas. The reactants in this reaction are
A. magnesium and magnesium chloride.
B. hydrochloric acid and hydrogen gas.
C. magnesium and hydrochloric acid.
D. magnesium chloride and hydrogen gas.

2. Matter that has a definite shape and a definite volume is
A. a liquid.
B. an element.
C. a solid.
D. a gas.

3. We know that air is a mixture and not a compound because
A. it can be heated to a higher temperature.
B. it can be compressed to a smaller volume.
C. it is colorless.
D. its composition can vary.

4. Matter can be defined as anything that
A. has weight.
B. has mass and volume.
C. is uniform throughout.
D. exhibits both chemical and physical properties.

5. Which of the following is best classified as a homogeneous mixture?
A. pizza
B. blood
C. hot tea
D. copper wire

6. A compound differs from a mixture in that a compound
A. contains only one element.
B. varies in chemical composition depending on the sample size.
C. has a definite composition by mass of the elements that the compound contains.
D. can be classified as either heterogeneous or homogeneous.

7. Which of the following is *not* a physical state of matter?
A. solid
B. gas
C. element
D. liquid

8. Three of the following must contain two or more kinds of atoms. Which one does *not* contain two or more kinds of atoms?
A. element
B. compound
C. homogeneous mixture
D. heterogeneous mixture

9. Which of the following symbols does *not* match the element name given?
A. Al, aluminum
B. Co, copper
C. K, potassium
D. P, phosphorus

SHORT ANSWER

10. Give three examples of mixtures, and tell whether each one is homogeneous or heterogeneous. Give three examples of compounds.

11. In trying to identify a sample of a pure substance, we observe the following properties. Tell whether each one is a chemical property or a physical property.
A. Its mass is 124.3 g.
B. It is a shiny solid at room temperature.
C. It is easily etched by nitric acid.
D. It melts when heated to 670°C.
E. It is 31.7 centimeters long.
F. It is a good heat conductor.
G. It burns in air.
H. It is a good conductor of electrical energy.

EXTENDED RESPONSE

12. Describe the difference between a chemical change and a physical change. Give one example of each kind of change.

13. Describe general properties of metals, nonmetals, and metalloids.

Test TIP Remember that if you can eliminate two of the four answer choices, your chances of choosing the correct answer choice will double.

Mixture Separation

OBJECTIVES

- *Observe* the chemical and physical properties of a mixture.
- *Relate* knowledge of chemical and physical properties to the task of purifying the mixture.
- *Analyze* the success of methods of purifying the mixture.

MATERIALS

- aluminum foil
- cotton balls
- distilled water
- filter funnels
- filter paper
- forceps
- magnet
- paper clips
- paper towels
- Petri dish
- pipets
- plastic forks
- plastic spoons
- plastic straws
- rubber stoppers
- sample of mixture and components (sand, iron filings, salt, poppy seeds)
- test tubes and rack
- tissue paper
- transparent tape
- wooden splints

BACKGROUND

The ability to separate and recover pure substances from mixtures is extremely important in scientific research and industry. Chemists need to work with pure substances, but naturally occurring materials are seldom pure. Often, differences in the physical properties of the components in a mixture provide the means for separating them. In this experiment, you will have an opportunity to design, develop, and implement your own procedure for separating a mixture. The mixture you will work with contains salt, sand, iron filings, and poppy seeds. All four substances are in dry, granular form.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Your task will be to plan and carry out the separation of a mixture. Before you can plan your experiment, you will need to investigate the properties of each component in the mixture. The properties will be used to design your mixture separation. Copy the data table on the following page in your lab notebook, and use it to record your observations.

PROCEDURE

1. Obtain separate samples of each of the four mixture components from your teacher. Use the equipment you have available to make observations of the components and determine their properties. You will need to run several tests with each substance, so don't use all of your sample

DATA TABLE				
Properties	Sand	Iron filings	Salt	Poppy seeds
Dissolves				
Floats				
Magnetic				
Other				

on the first test. Look for things like whether the substance is magnetic, whether it dissolves, or whether it floats. Record your observations in your data table.

2. Make a plan for what you will do to separate a mixture that includes the four components from step 1. Review your plan with your teacher.
3. Obtain a sample of the mixture from your teacher. Using the equipment you have available, run the procedure you have developed.

CLEANUP AND DISPOSAL

4. Clean your lab station. Clean all equipment, and return it to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or throw anything in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

ANALYSIS AND INTERPRETATION

1. **Evaluating Methods:** On a scale of 1 to 10, how successful were you in separating and recovering each of the four components: sand, salt, iron filings, and poppy seeds? Consider 1 to be the best and 10 to be the worst. Justify your ratings based on your observations.

CONCLUSIONS

1. **Evaluating Methods:** How did you decide on the order of your procedural steps? Would any order have worked?
2. **Designing Experiments:** If you could do the lab over again, what would you do differently? Be specific.
3. **Designing Experiments:** Name two materials or tools that weren't available that might have made your separation easier.
4. **Applying Ideas:** For each of the four components, describe a specific physical property that enabled you to separate the component from the rest of the mixture.

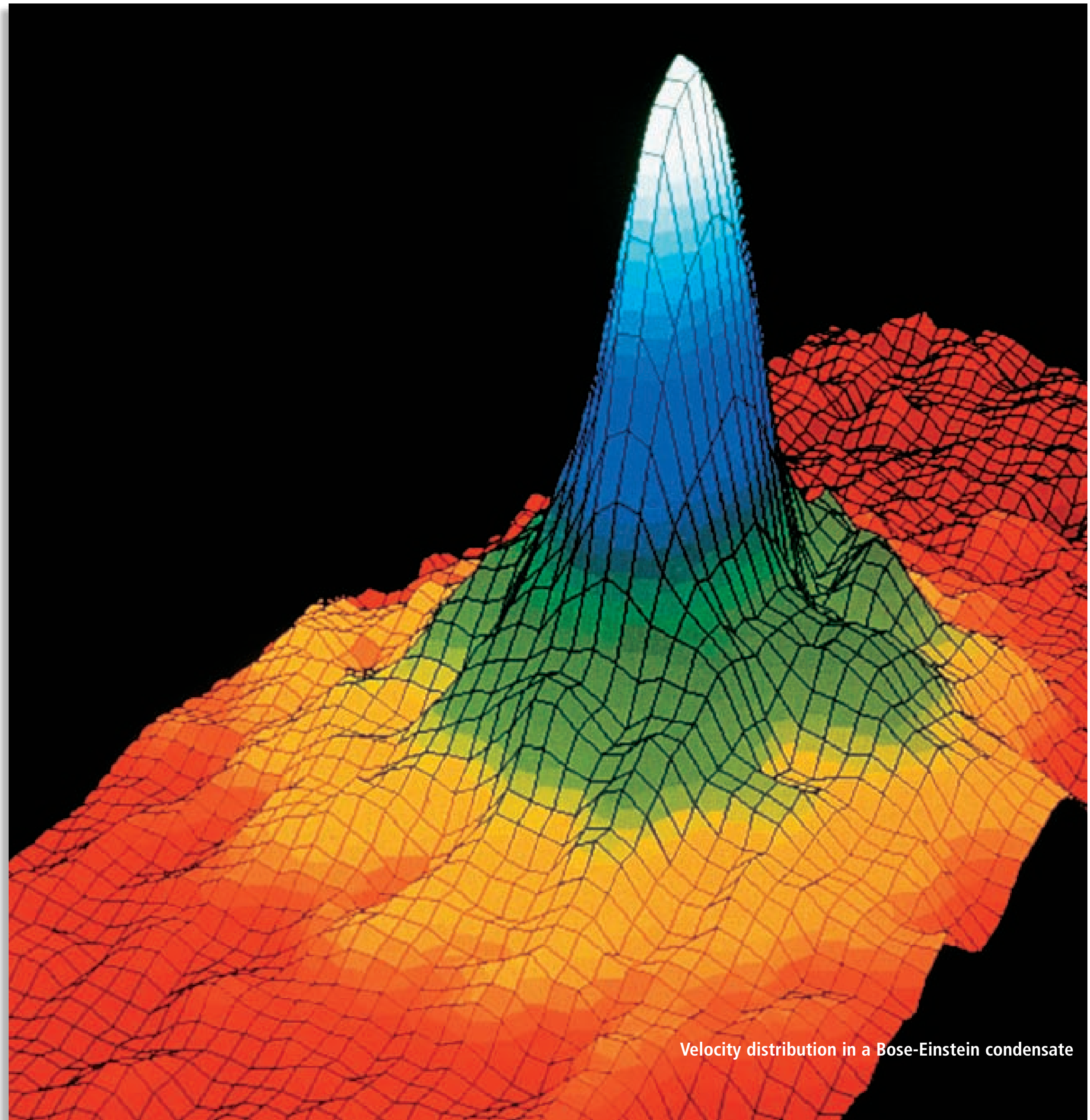
EXTENSIONS

1. **Evaluating Methods:** What methods could be used to determine the purity of each of your recovered components?
2. **Designing Experiments:** How could you separate each of the following two-part mixtures?
   a. aluminum filings and iron filings
   b. sand and gravel
   c. sand and finely ground polystyrene foam
   d. salt and sugar
   e. alcohol and water
   f. nitrogen and oxygen
3. **Designing Experiments:** One of the components of the mixture in this experiment is in a different physical state at the completion of this experiment that it was at the start. Which one? How would you convert that component back to its original state?

CHAPTER 2

Measurements and Calculations

Quantitative measurements are fundamental to chemistry.



Velocity distribution in a Bose-Einstein condensate

Scientific Method

SECTION 1

OBJECTIVES

- Describe the purpose of the scientific method.
- Distinguish between qualitative and quantitative observations.
- Describe the differences between hypotheses, theories, and models.

Sometimes progress in science comes about through accidental discoveries. Most scientific advances, however, result from carefully planned investigations. The process researchers use to carry out their investigations is often called the scientific method. *The* **scientific method** *is a logical approach to solving problems by observing and collecting data, formulating hypotheses, testing hypotheses, and formulating theories that are supported by data.*

Observing and Collecting Data

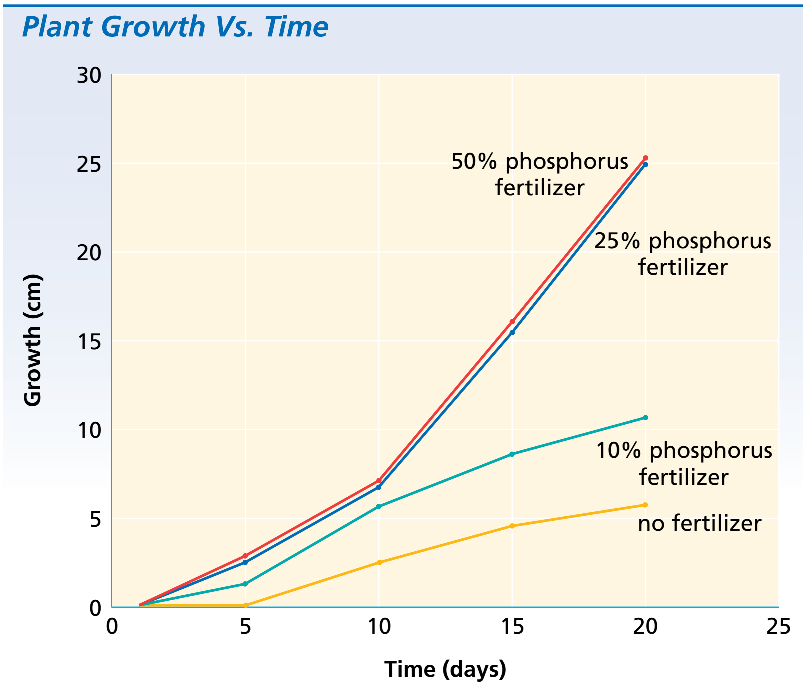
Observing is the use of the senses to obtain information. Observation often involves making measurements and collecting data. The data may be descriptive (qualitative) or numerical (quantitative) in nature. Numerical information, such as the fact that a sample of copper ore has a mass of 25.7 grams, is *quantitative*. Non-numerical information, such as the fact that the sky is blue, is *qualitative*.

Experimenting involves carrying out a procedure under controlled conditions to make observations and collect data. To learn more about matter, chemists study systems. *A* **system** *is a specific portion of matter in a given region of space that has been selected for study during an experiment or observation.* When you observe a reaction in a test tube, the test tube and its contents form a system.

SCILINKS.
www.scilinks.org
Topic: Scientific Methods
Code: HC61359

FIGURE 1 These students have designed an experiment to determine how to get the largest volume of popped corn from a fixed number of kernels. They think that the volume is likely to increase as the moisture in the kernels increases. Their experiment will involve soaking some kernels in water and observing whether the volume of the popped corn is greater than that of corn popped from kernels that have not been soaked.

FIGURE 2 A graph of data can show relationships between two variables. In this case the graph shows data collected during an experiment to determine the effect of phosphorus fertilizer compounds on plant growth. The following is one possible hypothesis: *If* phosphorus stimulates corn-plant growth, *then* corn plants treated with a soluble phosphorus compound should grow faster, under the same conditions, than corn plants that are not treated.



Formulating Hypotheses

As scientists examine and compare the data from their own experiments, they attempt to find relationships and patterns—in other words, they make generalizations based on the data. Generalizations are statements that apply to a range of information. To make generalizations, data are sometimes organized in tables and analyzed using statistics or other mathematical techniques, often with the aid of graphs and a computer.

Scientists use generalizations about the data to formulate a **hypothesis,** *or testable statement*. The hypothesis serves as a basis for making predictions and for carrying out further experiments. Hypotheses are often drafted as "if-then" statements. The "then" part of the hypothesis is a prediction that is the basis for testing by experiment. **Figure 2** shows data collected to test a hypothesis.

Testing Hypotheses

Testing a hypothesis requires experimentation that provides data to support or refute a hypothesis or theory. During testing, the experimental conditions that remain constant are called *controls,* and any condition that changes is called a *variable.* Any change observed is usually due to the effects of the variable. If testing reveals that the predictions were not correct, the hypothesis on which the predictions were based must be discarded or modified.

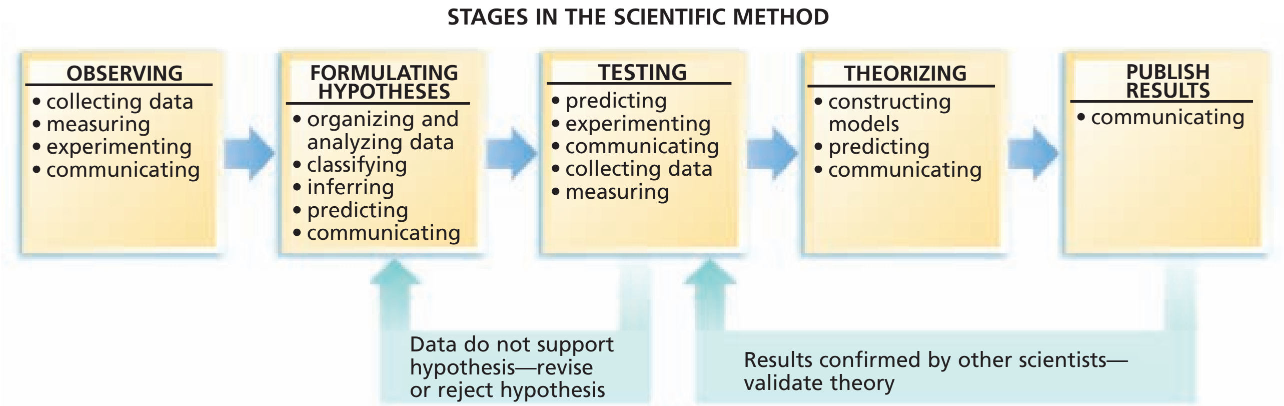


FIGURE 3 The scientific method is not a single, fixed process. Scientists may repeat steps many times before there is sufficient evidence to formulate a theory. You can see that each stage represents a number of different activities.

Theorizing

When the data from experiments show that the predictions of the hypothesis are successful, scientists typically try to explain the phenomena they are studying by constructing a model. *A* **model** *in science is more than a physical object; it is often an explanation of how phenomena occur and how data or events are related*. Models may be visual, verbal, or mathematical. One important model in chemistry is the atomic model of matter, which states that matter is composed of tiny particles called atoms.

If a model successfully explains many phenomena, it may become part of a theory. The atomic model is a part of the atomic theory, which you will study in Chapter 3. *A* **theory** *is a broad generalization that explains a body of facts or phenomena*. Theories are considered successful if they can predict the results of many new experiments. Examples of the important theories you will study in chemistry are kinetic-molecular theory and collision theory. **Figure 3** shows where theory fits in the scheme of the scientific method.

SECTION REVIEW

1. What is the scientific method?
2. Which of the following are quantitative?
   a. the liquid floats on water
   b. the metal is malleable
   c. the liquid has a temperature of 55.6°C
3. How do hypotheses and theories differ?
4. How are models related to theories and hypotheses?

Critical Thinking

5. **INTERPRETING CONCEPTS** Suppose you had to test how well two types of soap work. Describe your experiment by using the terms *control* and *variable.*

Chemistry in Action



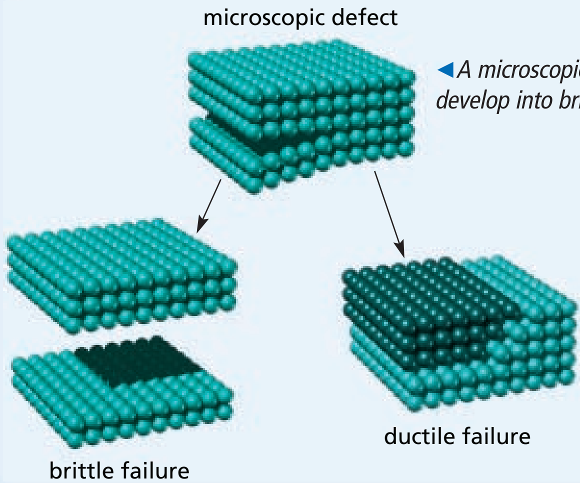
Breaking Up Is Easy To Do

It may seem obvious that chemistry is important in the making of materials, but chemistry is also vital to the study of how materials break. Everyday items have to be made to withstand various types of force and pressure or they cannot be used. For example, scientists and engineers work to ensure that highway bridges do not collapse.

When excessive force is applied to an object, the material that the object is made of will break. The object breaks because the force creates stress on the bonds between the atoms of the material and causes the bonds to break. This creates microscopic cracks in the material. When a material breaks, it is said to have undergone *failure.* Materials typically break in one of two ways: *ductile failure* and *brittle failure.* Both types of failure start with microscopic cracks in the material. However, the way a material eventually breaks depends how its atoms are organized.

Shattering glass undergoes brittle failure. Glass shatters when the bonds between the two layers of atoms that are along the initial crack break. This breakage causes the layers to pull apart, which separates the material into pieces. This type of failure is common in materials that do not have a very orderly arrangement of atoms.

When a car bumper crumples, ductile failure happens. This type of failure tends to happen in materials such as metals, that have a regular, ordered arrangement of atoms. This arrangement of atoms is known as a *crystal structure.* Ductile failure happens when the bonds in the material break across many layers of atoms that are not in the same plane as the original crack. Rather than splitting apart, the layers slip past each other into new positions. The atoms form new chemical bonds, between them and the material stays in one piece; only the shape has changed.



◀ *A microscopic crack in a material can develop into brittle or ductile failure.*

In addition to the type of material influencing breakage, the quality of the material also influences breakage. All objects contain microscopic defects, such as bubbles in plastic pieces. A material will tend to undergo failure at its defect sites first. Careful fabrication procedures can minimize, but not completely eliminate, defects in materials.

Even though materials are designed to withstand a certain amount of force, the normal wear and tear that materials experience over their lifetimes creates defects in the material. This process is referred to as *fatigue.* If fatigue were to go undetected, the microscopic cracks that form could then undergo brittle or ductile failure. It would be catastrophic if the materials in certain products, such as airplane parts, failed. To avoid such a failure, people monitor materials that are exposed to constant stress for signs of fatigue. The defects in the metal parts of airplanes can be detected with nondestructive techniques, such as electromagnetic analysis.

Questions

1. Can you name some ways in which metal or plastic parts might obtain defects caused by chemical reactions?
2. Does a ceramic dinner plate undergo brittle or ductile failure when it is dropped and breaks?

Units of Measurement

SECTION 2

Objectives

- Distinguish between a quantity, a unit, and a measurement standard.
- Name and use SI units for length, mass, time, volume, and density.
- Distinguish between mass and weight.
- Perform density calculations.
- Transform a statement of equality into a conversion factor.

Measurements are quantitative information. A measurement is more than just a number, even in everyday life. Suppose a chef were to write a recipe listing quantities such as 1 salt, 3 sugar, and 2 flour. The cooks could not use the recipe without more information. They would need to know whether the number 3 represented teaspoons, tablespoons, cups, ounces, grams, or some other unit for sugar.

Measurements *represent* quantities. *A* **quantity** *is something that has magnitude, size, or amount.* A quantity is not the same as a measurement. For example, the quantity represented by a teaspoon is volume. The teaspoon is a unit of measurement, while volume is a quantity. A teaspoon is a measurement standard in this country. Units of measurement compare what is to be measured with a previously defined size. Nearly every measurement is a number plus a unit. The choice of unit depends on the quantity being measured.

Many centuries ago, people sometimes marked off distances in the number of foot lengths it took to cover the distance. But this system was unsatisfactory because the number of foot lengths used to express a distance varied with the size of the measurer's foot. Once there was agreement on a standard for foot length, confusion as to the real length was eliminated. It no longer mattered who made the measurement, as long as the standard measuring unit was correctly applied.

SI Measurement

Scientists all over the world have agreed on a single measurement system called *Le Système International d'Unités*, abbreviated **SI.** This system was adopted in 1960 by the General Conference on Weights and Measures. SI now has seven base units, and most other units are derived from these seven. Some non-SI units are still commonly used by chemists and are also used in this book.

SI units are defined in terms of standards of measurement. The standards are objects or natural phenomena that are of constant value, easy to preserve and reproduce, and practical in size. International organizations monitor the defining process. In the United States, the National Institute of Standards and Technology (NIST) plays the main role in maintaining standards and setting style conventions. For example, numbers are written in a form that is agreed upon internationally. The number seventy-five thousand is written 75 000, not 75,000, because the comma is used in other countries to represent a decimal point.

TABLE 1 SI Base Units

Quantity	Quantity symbol	Unit name	Unit abbreviation	Defined standard
Length	l	meter	m	the length of the path traveled by light in a vacuum during a time interval of 1/299 792 458 of a second
Mass	m	kilogram	kg	the unit of mass equal to the mass of the international prototype of the kilogram
Time	t	second	s	the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom
Temperature	T	kelvin	K	the fraction 1/273.16 of the thermodynamic temperature of the triple point of water
Amount of substance	n	mole	mol	the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12
Electric current	I	ampere	A	the constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross section, and placed 1 meter apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per meter of length
Luminous intensity	I_v	candela	cd	the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian

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Topic: SI Units
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SI Base Units

The seven SI base units and their standard abbreviated symbols are listed in **Table 1.** All the other SI units can be derived from the fundamental units.

Prefixes added to the names of SI base units are used to represent quantities that are larger or smaller than the base units. **Table 2** lists SI prefixes using units of length as examples. For example, the prefix *centi-*, abbreviated c, represents an exponential factor of 10^{-2}, which equals 1/100. Thus, 1 centimeter, 1 cm, equals 0.01 m, or 1/100 of a meter.

Mass

As you learned in Chapter 1, mass is a measure of the quantity of matter. The SI standard unit for mass is the kilogram. The standard for mass defined in **Table 1** is used to calibrate balances all over the world.

TABLE 2 SI Prefixes

Prefix	Unit abbreviation	Exponential factor	Meaning	Example
tera	T	10^{12}	1 000 000 000 000	1 terameter (Tm) = 1×10^{12} m
giga	G	10^{9}	1 000 000 000	1 gigameter (Gm) = 1×10^{9} m
mega	M	10^{6}	1 000 000	1 megameter (Mm) = 1×10^{6} m
kilo	k	10^{3}	1000	1 kilometer (km) = 1000 m
hecto	h	10^{2}	100	1 hectometer (hm) = 100 m
deka	da	10^{1}	10	1 dekameter (dam) = 10 m
		$\mathbf{10^{0}}$	**1**	**1 meter (m)**
deci	d	10^{-1}	1/10	1 decimeter (dm) = 0.1 m
centi	c	10^{-2}	1/100	1 centimeter (cm) = 0.01 m
milli	m	10^{-3}	1/1000	1 millimeter (mm) = 0.001 m
micro	μ	10^{-6}	1/1 000 000	1 micrometer (μm) = 1×10^{-6} m
nano	n	10^{-9}	1/1 000 000 000	1 nanometer (nm) = 1×10^{-9} m
pico	p	10^{-12}	1/1 000 000 000 000	1 picometer (pm) = 1×10^{-12} m
femto	f	10^{-15}	1/1 000 000 000 000 000	1 femtometer (fm) = 1×10^{-15} m
atto	a	10^{-18}	1/1 000 000 000 000 000 000	1 attometer (am) = 1×10^{-18} m

The mass of a typical textbook is about 1 kg. The gram, g, which is 1/1000 of a kilogram, is more useful for measuring masses of small objects, such as flasks and beakers. For even smaller objects, such as tiny quantities of chemicals, the milligram, mg, is often used. One milligram is 1/1000 of a gram, or 1/1 000 000 of a kilogram.

Mass is often confused with weight because people often express the weight of an object in grams. Mass is determined by comparing the mass of an object with a set of standard masses that are part of the balance. **Weight** *is a measure of the gravitational pull on matter.* Unlike weight, mass does not depend on gravity. Mass is measured on instruments such as a balance, and weight is typically measured on a spring scale. Taking weight measurements involves reading the amount that an object pulls down on a spring. As the force of Earth's gravity on an object increases, the object's weight increases. The weight of an object on the moon is about one-sixth of its weight on Earth.

Length

The SI standard unit for length is the meter. A distance of 1 m is about the width of an average doorway. To express longer distances, the kilometer, km, is used. One kilometer equals 1000 m. Road signs in the United States sometimes show distances in kilometers as well as miles. The kilometer is the unit used to express highway distances in most other countries of the world. To express shorter distances, the centimeter

CROSS-DISCIPLINARY

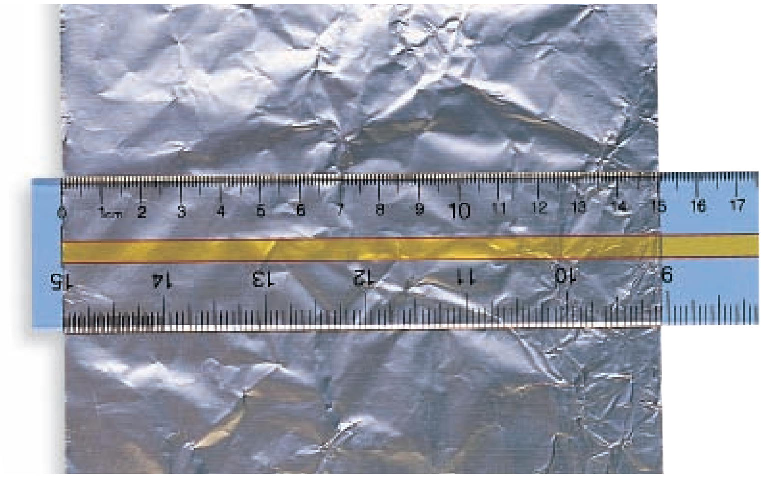
Some Handy Comparisons of Units

To become comfortable with units in the SI system, try relating some common measurements to your experience.

A meter stick is a little longer than a yardstick. A millimeter is about the diameter of a paper clip wire, and a centimeter is a little more than the width of a paper clip.

One gram is about the mass of a paper clip. A kilogram is about 2.2 pounds (think of two pounds plus one stick of butter). And there are about five milliliters in a teaspoon.

FIGURE 4 The meter is the SI unit of length, but the centimeter is often used to measure smaller distances. What is the length in cm of the rectangular piece of aluminum foil shown?



is often used. From **Table 2,** you can see that one centimeter equals 1/100 of a meter. The width of this book is just over 20 cm.

Derived SI Units

Many SI units are combinations of the quantities shown in **Table 1.** *Combinations of SI base units form* **derived units.** Some derived units are shown in **Table 3.**

Derived units are produced by multiplying or dividing standard units. For example, area, a derived unit, is length times width. If both length and width are expressed in meters, the area unit equals meters times meters, or square meters, abbreviated m^2. The last column of

TABLE 3 ***Derived SI Units***

Quantity	Quantity symbol	Unit	Unit abbreviation	Derivation
Area	A	square meter	m^2	length × width
Volume	V	cubic meter	m^3	length × width × height
Density	D	kilograms per cubic meter	$\frac{kg}{m^3}$	$\frac{\text{mass}}{\text{volume}}$
Molar mass	M	kilograms per mole	$\frac{kg}{mol}$	$\frac{\text{mass}}{\text{amount of substance}}$
Molar volume	V_m	cubic meters per mole	$\frac{m^3}{mol}$	$\frac{\text{volume}}{\text{amount of substance}}$
Energy	E	joule	J	force × length

Table 3 shows the combination of fundamental units used to obtain derived units.

Some combination units are given their own names. For example, pressure expressed in base units is the following.

$$\text{kg/m}\bullet\text{s}^2$$

The name *pascal,* Pa, is given to this combination. You will learn more about pressure in Chapter 11. Prefixes can also be added to express derived units. Area can be expressed in cm^2, square centimeters, or mm^2, square millimeters.

FIGURE 5 The speed that registers on a speedometer represents distance traveled per hour and is expressed in the derived units kilometers per hour or miles per hour.

Volume

Volume *is the amount of space occupied by an object.* The derived SI unit of volume is cubic meters, m^3. One cubic meter is equal to the volume of a cube whose edges are 1 m long. Such a large unit is inconvenient for expressing the volume of materials in a chemistry laboratory. Instead, a smaller unit, the cubic centimeter, cm^3, is often used. There are 100 centimeters in a meter, so a cubic meter contains 1 000 000 cm^3.

$$1\ \text{m}^3 \times \frac{100\ \text{cm}}{1\ \text{m}} \times \frac{100\ \text{cm}}{1\ \text{m}} \times \frac{100\ \text{cm}}{1\ \text{m}} = 1\ 000\ 000\ \text{cm}^3$$

When chemists measure the volumes of liquids and gases, they often use a non-SI unit called the liter. The liter is equivalent to one cubic decimeter. Thus, a liter, L, is also equivalent to 1000 cm^3. Another non-SI unit, the milliliter, mL, is used for smaller volumes. There are 1000 mL in 1 L. Because there are also 1000 cm^3 in a liter, the two units—milliliter and cubic centimeter—are interchangeable.

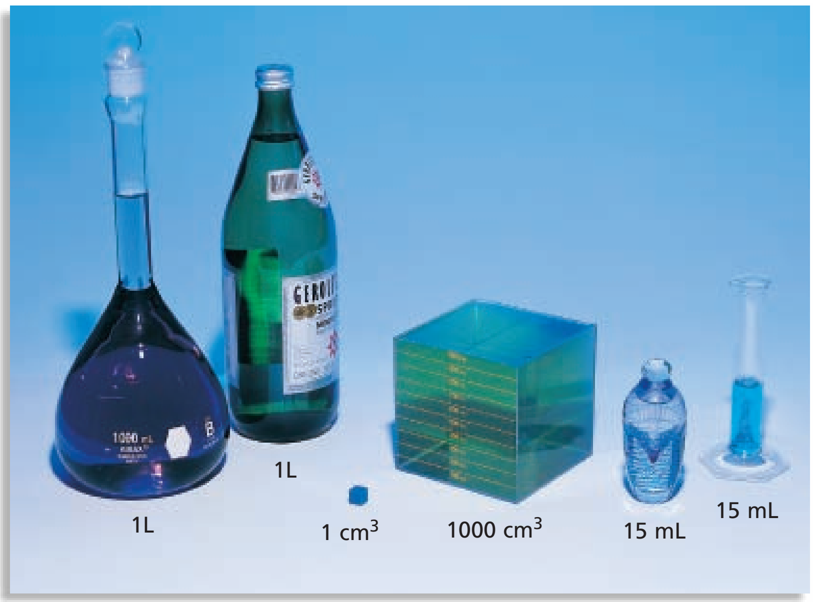


FIGURE 6 The relationships between various volumes are shown here. One liter contains 1000 mL of liquid, and 1 mL is equivalent to 1 cm^3. A small perfume bottle contains about 15 mL of liquid. The volumetric flask (far left) and graduated cylinder (far right) are used for measuring liquid volumes in the lab.

FIGURE 7 Density is the ratio of mass to volume. Both water and copper shot float on mercury because mercury is more dense.

Density

An object made of cork feels lighter than a lead object of the same size. What you are actually comparing in such cases is how massive objects are compared with their size. This property is called density. **Density** *is the ratio of mass to volume, or mass divided by volume*. Mathematically, the relationship for density can be written in the following way.

$$density = \frac{mass}{volume} \text{ or } D = \frac{m}{V}$$

The quantity m is mass, V is volume, and D is density.

The SI unit for density is derived from the base units for mass and volume—the kilogram and the cubic meter, respectively—and can be expressed as kilograms per cubic meter, kg/m^3. This unit is inconveniently large for the density measurements you will make in the laboratory. You will often see density expressed in grams per cubic centimeter, g/cm^3, or grams per milliliter, g/mL. The densities of gases are generally reported either in kilograms per cubic meter, kg/m^3, or in grams per liter, g/L.

Density is a characteristic physical property of a substance. It does not depend on the size of the sample because as the sample's mass increases, its volume increases proportionately, and the ratio of mass to volume is constant. Therefore, density can be used as one property to help identify a substance. **Table 4** shows the densities of some common materials. As you can see, cork has a density of only 0.24 g/cm^3, which is less than the density of liquid water. Because cork is less dense than water, it floats on water. Lead, on the other hand, has a density of 11.35 g/cm^3. The density of lead is greater than that of water, so lead sinks in water.

Note that **Table 4** specifies the temperatures at which the densities were measured. That is because density varies with temperature. Most objects expand as temperature increases, thereby increasing in volume. Because density is mass divided by volume, density usually decreases with increasing temperature.

TABLE 4 ***Densities of Some Familiar Materials***

Solids	Density at 20°C (g/cm^3)	Liquids	Density at 20°C (g/mL)
cork	0.24*	gasoline	0.67*
butter	0.86	ethyl alcohol	0.791
ice	0.92†	kerosene	0.82
sucrose	1.59	turpentine	0.87
bone	1.85*	water	0.998
diamond	3.26*	sea water	1.025**
copper	8.92	milk	1.031*
lead	11.35	mercury	13.6

† measured at 0°C
* typical density
** measured at 15°C

QuickLAB

Wear safety goggles and an apron.

Density of Pennies

Materials
- balance
- 100 mL graduated cylinder
- 40 pennies dated before 1982
- 40 pennies dated after 1982
- water

Procedure

1. Using the balance, determine the mass of the 40 pennies minted prior to 1982. Repeat this measurement two more times. Average the results of the three trials to determine the average mass of the pennies.
2. Repeat step 1 with the 40 pennies minted after 1982.
3. Pour about 50 mL of water into the 100 mL graduated cylinder. Record the exact volume of the water. Add the 40 pennies minted before 1982. CAUTION: Add the pennies carefully so that no water is splashed out of the cylinder. Record the exact volume of the water and pennies. Repeat this process two more times. Determine the volume of the pennies for each trial. Average the results of those trials to determine the average volume of the pennies.
4. Repeat step 3 with the 40 pennies minted after 1982.
5. Review your data for any large differences between trials that could increase the error of your results. Repeat those measurements.
6. Use the average volume and average mass to calculate the average density for each group of pennies.
7. Compare the calculated average densities with the density of the copper listed in Table 4.

Discussion

1. Why is it best to use the results of three trials rather than a single trial for determining the density?
2. How did the densities of the two groups of pennies compare? How do you account for any difference?
3. Use the results of this investigation to formulate a hypothesis about the composition of the two groups of pennies. How could you test your hypothesis?

SAMPLE PROBLEM A

A sample of aluminum metal has a mass of 8.4 g. The volume of the sample is 3.1 cm^3. Calculate the density of aluminum.

SOLUTION

Given: mass (m) = 8.4 g
volume (V) = 3.1 cm^3

Unknown: density (D)

$$density = \frac{mass}{volume} = \frac{8.4\ \text{g}}{3.1\ \text{cm}^3} = 2.7\ \text{g/cm}^3$$

PRACTICE *Answers in Appendix E*

1. What is the density of a block of marble that occupies 310. cm^3 and has a mass of 853 g?
2. Diamond has a density of 3.26 g/cm^3. What is the mass of a diamond that has a volume of 0.351 cm^3?
3. What is the volume of a sample of liquid mercury that has a mass of 76.2 g, given that the density of mercury is 13.6 g/mL?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate density.

Keyword: HC6MEAX

Conversion Factors

A **conversion factor** *is a ratio derived from the equality between two different units that can be used to convert from one unit to the other.* For example, suppose you want to know how many quarters there are in a certain number of dollars. To figure out the answer, you need to know how quarters and dollars are related. There are four quarters per dollar and one dollar for every four quarters. Those facts can be expressed as ratios in four conversion factors.

$$\frac{4 \text{ quarters}}{1 \text{ dollar}} = 1 \qquad \frac{1 \text{ dollar}}{4 \text{ quarters}} = 1 \qquad \frac{0.25 \text{ dollar}}{1 \text{ quarter}} = 1 \qquad \frac{1 \text{ quarter}}{0.25 \text{ dollar}} = 1$$

Notice that each conversion factor equals 1. That is because the two quantities divided in any conversion factor are equivalent to each other—as in this case, where 4 quarters equal 1 dollar. Because conversion factors are equal to 1, they can be multiplied by other factors in equations without changing the validity of the equations. You can use conversion factors to solve problems through dimensional analysis. **Dimensional analysis** *is a mathematical technique that allows you to use units to solve problems involving measurements.* When you want to use a conversion factor to change a unit in a problem, you can set up the problem in the following way.

$$\text{quantity sought} = \text{quantity given} \times \text{conversion factor}$$

For example, to determine the number of quarters in 12 dollars, you would carry out the unit conversion that allows you to change from dollars to quarters.

$$\text{number of quarters} = 12 \text{ dollars} \times \text{conversion factor}$$

Next you would have to decide which conversion factor gives you an answer in the desired unit. In this case, you have dollars and you want quarters. To eliminate dollars, you must divide the quantity by dollars. Therefore, the conversion factor in this case must have dollars in the denominator and quarters in the numerator. That factor is 4 quarters/1 dollar. Thus, you would set up the calculation as follows.

$$? \text{ quarters} = 12 \text{ dollars} \times \text{conversion factor}$$

$$= 12 \cancel{\text{dollars}} \times \frac{4 \text{ quarters}}{1 \cancel{\text{dollar}}} = 48 \text{ quarters}$$

Notice that the dollars have divided out, leaving an answer in the desired unit—quarters.

Suppose you had guessed wrong and used 1 dollar/4 quarters when choosing which of the two conversion factors to use. You would have an answer with entirely inappropriate units.

$$? \text{ quarters} = 12 \text{ dollars} \times \frac{1 \text{ dollar}}{4 \text{ quarters}} = \frac{3 \text{ dollars}^2}{\text{quarter}}$$

It is always best to begin with an idea of the units you will need in your final answer. When working through the Sample Problems, keep track of the units needed for the unknown quantity. Check your final answer against what you've written as the unknown quantity.

Deriving Conversion Factors

You can derive conversion factors if you know the relationship between the unit you have and the unit you want. For example, from the fact that *deci-* means "1/10," you know that there is 1/10 of a meter per decimeter and that each meter must have 10 decimeters. Thus, from the equality (1 m = 10 dm), you can write the following conversion factors relating meters and decimeters. In this book, when there is no digit shown in the denominator, you can assume the value is 1.

$$\frac{1 \text{ m}}{10 \text{ dm}} \text{ and } \frac{0.1 \text{ m}}{\text{dm}} \text{ and } \frac{10 \text{ dm}}{\text{m}}$$

The following sample problem illustrates an example of deriving conversion factors to make a unit conversion.

SAMPLE PROBLEM B

Express a mass of 5.712 grams in milligrams and in kilograms.

SOLUTION

Given: 5.712 g

Unknown: mass in mg and kg

The expression that relates grams to milligrams is

$$1 \text{ g} = 1000 \text{ mg}$$

The possible conversion factors that can be written from this expression are

$$\frac{1000 \text{ mg}}{\text{g}} \text{ and } \frac{1 \text{ g}}{1000 \text{ mg}}$$

To derive an answer in mg, you'll need to multiply 5.712 g by 1000 mg/g.

$$5.712\ \cancel{g} \times \frac{1000\ \text{mg}}{\cancel{g}} = 5712\ \text{mg}$$

This answer makes sense because milligrams is a smaller unit than grams and, therefore, there should be more of them.

The kilogram problem is solved similarly.

$$1\ \text{kg} = 1000\ \text{g}$$

Conversion factors representing this expression are

$$\frac{1\ \text{kg}}{1000\ \text{g}} \text{ and } \frac{1000\ \text{g}}{\text{kg}}$$

To derive an answer in kg, you'll need to multiply 5.712 g by 1 kg/1000 g.

$$5.712\ \cancel{g} \times \frac{1\ \text{kg}}{1000\ \cancel{g}} = 0.005712\ \text{kg}$$

The answer makes sense because kilograms is a larger unit than grams and, therefore, there should be fewer of them.

PRACTICE

Answers in Appendix E

1. Express a length of 16.45 m in centimeters and in kilometers.
2. Express a mass of 0.014 mg in grams.

extension

Go to **go.hrw.com** for more practice problems that ask you to perform unit conversions.

Keyword: HC6MEAX

SECTION REVIEW

1. Why are standards needed for measured quantities?
2. Label each of the following measurements by the quantity each represents. For instance, a measurement of 10.6 kg/m^3 represents density.
   - **a.** 5.0 g/mL
   - **b.** 37 s
   - **c.** 47 J
   - **d.** 39.56 g
   - **e.** 25.3 cm^3
   - **f.** 325 ms
   - **g.** 500 m^2
   - **h.** 30.23 mL
   - **i.** 2.7 mg
   - **j.** 0.005 L
3. Complete the following conversions.
   - **a.** 10.5 g = _____ kg
   - **b.** 1.57 km = _____ m
   - **c.** 3.54 μg = _____ g
   - **d.** 3.5 mol = _____ μmol
   - **e.** 1.2 L = _____ mL
   - **f.** 358 cm^3 = _____ m^3
   - **g.** 548.6 mL = _____ cm^3
4. Write conversion factors for each equality.
   - **a.** 1 m^3 = 1 000 000 cm^3
   - **b.** 1 in. = 2.54 cm
   - **c.** 1 μg = 0.000 001 g
   - **d.** 1 Mm = 1 000 000 m
5. **a.** What is the density of an 84.7 g sample of an unknown substance if the sample occupies 49.6 cm^3?
   b. What volume would be occupied by 7.75 g of this same substance?

Critical Thinking

6. **INFERRING CONCLUSIONS** A student converts grams to milligrams by multiplying by the conversion factor $\frac{1\ \text{g}}{1000\ \text{mg}}$. Is the student performing this calculation correctly?

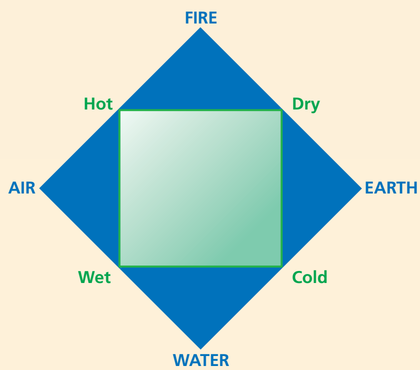
HISTORICAL CHEMISTRY

Classical Ideas About Matter

The Greeks were among the many ancient cultures that sought to understand the nature of matter. One group of Greek philosophers, called the *atomists,* believed that matter could be broken down into pieces of a minute size. These pieces, called *atoms* or *atomos* which means "indivisible," possessed intrinsic, unchanging qualities. Another group of Greeks believed that matter could be divided an infinite number of times and could be changed from one type of matter into another.

Between 500 and 300 BCE, the Greek philosophers Leucippus and Democritus formulated the ideas that the atomists held. Leucippus and Democritus believed that all atoms were essentially the same but that the properties of all substances arose from the unique characteristics of their atoms. For example, solids, such as most metals, were thought to have uneven, jagged atoms. Because the atoms were rough, they could stick together and form solids. Similarly, water was thought to have atoms with smooth surfaces, which would allow the atoms to flow past one another. Though atomists did not have the same ideas about matter that we have today, they did believe that atoms were constantly in motion, even in objects that appeared to be solid.

Some Greek philosophers who studied matter between 700 and 300 BCE described matter in a way that differed from the way atomists described it. They attempted to identify and describe a fundamental substance from which all other matter was formed. Thales of Miletus (640–546 BCE) was among the first to suggest the existence of a basic element. He chose water, which exists as liquid, ice, and steam. He interpreted water's changeability to mean that water could transform into any other substance. Other philosophers suggested that the basic element was air or fire. Empedokles (ca. 490–ca. 430 BCE) focused on four elements: earth, air, fire, and water. He thought that these elements combined in various proportions to make all known matter.



▲ *This diagram shows Aristotle's belief about the relationship between the basic elements and properties.*

Aristotle (384–322 BCE), a student of Plato, elaborated on the earlier ideas about elements. He argued that in addition to the four elements that make up all matter, there were four basic properties: hot, cold, wet, and dry. In Aristotle's view, the four elements could each have two of the basic properties. For example, water was wet and cold, while air was wet and hot. He thought that one element could change into another element if its properties were changed.

For more than 2,000 years, Aristotle's classical ideas dominated scientific thought. His ideas were based on philosophical arguments, not on the the scientific process. It was not until the 1700s that the existence of atoms was shown experimentally and that the incredible intuition of the atomists was realized.

Questions

1. In Aristotle's system of elements, fire opposes water. Why do you think that he chose this relationship?
2. Use the ideas of the atomists to describe the atoms of the physical phases of matter—solid, liquid, and gas.

SECTION 3

OBJECTIVES

- Distinguish between accuracy and precision.
- Determine the number of significant figures in measurements.
- Perform mathematical operations involving significant figures.
- Convert measurements into scientific notation.
- Distinguish between inversely and directly proportional relationships.

Using Scientific Measurements

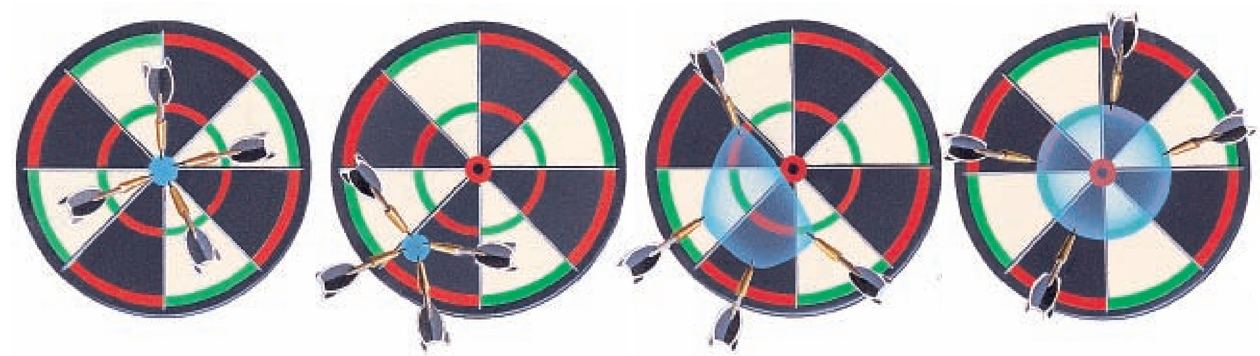
If you have ever measured something several times, you know that the results can vary. In science, for a reported measurement to be useful, there must be some indication of its reliability or uncertainty.

Accuracy and Precision

The terms *accuracy* and *precision* mean the same thing to most people. However, in science their meanings are quite distinct. **Accuracy** *refers to the closeness of measurements to the correct or accepted value of the quantity measured*. **Precision** *refers to the closeness of a set of measurements of the same quantity made in the same way*. Thus, measured values that are accurate are close to the accepted value. Measured values that are precise are close to one another but not necessarily close to the accepted value.

Figure 8 can help you visualize the difference between precision and accuracy. A set of darts thrown separately at a dartboard may land in various positions, relative to the bull's-eye and to one another. The

FIGURE 8 The sizes and locations of the areas covered by thrown darts illustrate the difference between precision and accuracy.



(a) Darts within small area = High precision; Area centered on bull's-eye = High accuracy

(b) Darts within small area = High precision; Area far from bull's-eye = Low accuracy

(c) Darts within large area = Low precision; Area far from bull's-eye = Low accuracy

(d) Darts within large area = Low precision; Area centered around bull's-eye = High accuracy (on average) closer the darts land to the bull's-eye, the more accurately they were thrown. The closer they land to one another, the more precisely they were thrown. Thus, the set of results shown in **Figure 8a** is both accurate and precise because the darts are close to the bull's-eye and close to each other. In **Figure 8b,** the set of results is inaccurate but precise because the darts are far from the bull's-eye but close to each other. In **Figure 8c,** the set of results is both inaccurate and imprecise because the darts are far from the bull's-eye and far from each other. Notice also that the darts are not evenly distributed around the bull's-eye, so the set, even considered on average, is inaccurate. In **Figure 8d,** the set on average is accurate compared with the third case, but it is imprecise. That is because the darts are distributed evenly around the bull's-eye but are far from each other.

extension

Chemistry in Action

Go to **go.hrw.com** for a full-length article on using measurements to determine a car's pollution rating.

Keyword: HC6MEAX

Percentage Error

The accuracy of an individual value or of an average experimental value can be compared quantitatively with the correct or accepted value by calculating the percentage error. **Percentage error** *is calculated by subtracting the accepted value from the experimental value, dividing the difference by the accepted value, and then multiplying by 100.*

$$\textit{Percentage error} = \frac{\textit{Value}_{\textit{experimental}} - \textit{Value}_{\textit{accepted}}}{\textit{Value}_{\textit{accepted}}} \times 100$$

Percentage error has a negative value if the accepted value is greater than the experimental value. It has a positive value if the accepted value is less than the experimental value. The following sample problem illustrates the concept of percentage error.

SAMPLE PROBLEM C

A student measures the mass and volume of a substance and calculates its density as 1.40 g/mL. The correct, or accepted, value of the density is 1.30 g/mL. What is the percentage error of the student's measurement?

SOLUTION

$$\textit{Percentage error} = \frac{\textit{Value}_{\textit{experimental}} - \textit{Value}_{\textit{accepted}}}{\textit{Value}_{\textit{accepted}}} \times 100$$

$$= \frac{1.40\ \cancel{\text{g/mL}} - 1.30\ \cancel{\text{g/mL}}}{1.30\ \cancel{\text{g/mL}}} \times 100 = 7.7\%$$

PRACTICE

Answers in Appendix E

1. What is the percentage error for a mass measurement of 17.7 g, given that the correct value is 21.2 g?
2. A volume is measured experimentally as 4.26 mL. What is the percentage error, given that the correct value is 4.15 mL?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate percentage error.

Keyword: HC6MEAX

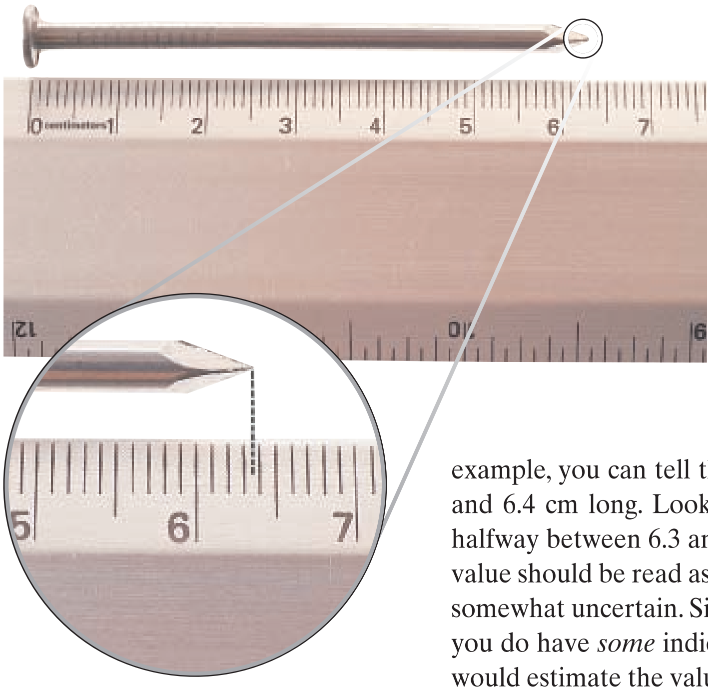


FIGURE 9 What value should be recorded for the length of this nail?

Error in Measurement

Some error or uncertainty always exists in any measurement. The skill of the measurer places limits on the reliability of results. The conditions of measurement also affect the outcome. The measuring instruments themselves place limitations on precision. Some balances can be read more precisely than others. The same is true of rulers, graduated cylinders, and other measuring devices.

When you use a properly calibrated measuring device, you can be almost certain of a particular number of digits in a reading. For example, you can tell that the nail in **Figure 9** is definitely between 6.3 and 6.4 cm long. Looking more closely, you can see that the value is halfway between 6.3 and 6.4 cm. However, it is hard to tell whether the value should be read as 6.35 cm or 6.36 cm. The hundredths place is thus somewhat uncertain. Simply leaving it out would be misleading because you do have *some* indication of the value's likely range. Therefore, you would estimate the value to the final questionable digit, perhaps reporting the length of the nail as 6.36 cm. You might include a plus-or-minus value to express the range, for example, 6.36 cm ± 0.01 cm.

Significant Figures

In science, measured values are reported in terms of significant figures. **Significant figures** *in a measurement consist of all the digits known with certainty plus one final digit, which is somewhat uncertain or is estimated*. For example, in the reported nail length of 6.36 cm discussed above, the last digit, 6, is uncertain. All the digits, including the uncertain one, are significant, however. All contain information and are included in the reported value. Thus, the term *significant* does not mean *certain*. In any correctly reported measured value, the final digit is significant but not certain. Insignificant digits are never reported. As a chemistry student, you will need to use and recognize significant figures when you work with measured quantities and report your results, and when you evaluate measurements reported by others.

Determining the Number of Significant Figures

When you look at a measured quantity, you need to determine which digits are significant. That process is very easy if the number has no zeros because all the digits shown are significant. For example, in a number reported as 3.95, all three digits are significant. The significance of zeros in a number depends on their location, however. You need to learn and follow several rules involving zeros. After you have studied the rules in **Table 5,** use them to express the answers in the sample problem that follows.

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Topic: Significant Figures
Code: HC61392

TABLE 5 *Rules for Determining Significant Zeros*

Rule	Examples
1. Zeros appearing between nonzero digits are significant.	a. 40.7 L has three significant figures. b. 87 009 km has five significant figures.
2. Zeros appearing in front of all nonzero digits are not significant.	a. 0.095 897 m has five significant figures. b. 0.0009 kg has one significant figure.
3. Zeros at the end of a number and to the right of a decimal point are significant.	a. 85.00 g has four significant figures. b. 9.000 000 000 mm has 10 significant figures.
4. Zeros at the end of a number but to the left of a decimal point may or may not be significant. If a zero has not been measured or estimated but is just a placeholder, it is not significant. A decimal point placed after zeros indicates that they are significant.	a. 2000 m may contain from one to four significant figures, depending on how many zeros are placeholders. **For measurements given in this text, assume that 2000 m has one significant figure.** b. 2000. m contains four significant figures, indicated by the presence of the decimal point.

SAMPLE PROBLEM D

For more help, go to the Math Tutor at the end of Chapter 1.

How many significant figures are in each of the following measurements?

a. 28.6 g
b. 3440. cm
c. 910 m
d. 0.046 04 L
e. 0.006 700 0 kg

SOLUTION Determine the number of significant figures in each measurement using the rules listed in **Table 5.**

a. 28.6 g
There are no zeros, so all three digits are significant.

b. 3440. cm
By rule 4, the zero is significant because it is immediately followed by a decimal point; there are 4 significant figures.

c. 910 m
By rule 4, the zero is not significant; there are 2 significant figures.

d. 0.046 04 L
By rule 2, the first two zeros are not significant; by rule 1, the third zero is significant; there are 4 significant figures.

e. 0.006 700 0 kg
By rule 2, the first three zeros are not significant; by rule 3, the last three zeros are significant; there are 5 significant figures.

PRACTICE *Answers in Appendix E*

1. Determine the number of significant figures in each of the following.
   a. 804.05 g
   b. 0.014 403 0 km
   c. 1002 m
   d. 400 mL
   e. 30 000. cm
   f. 0.000 625 000 kg

2. Suppose the value "seven thousand centimeters" is reported to you. How should the number be expressed if it is intended to contain the following?
   a. 1 significant figure
   b. 4 significant figures
   c. 6 significant figures

extension

Go to **go.hrw.com** for more practice problems that ask you to determine significant figures.

Keyword: HC6MEAX

Rounding

When you perform calculations involving measurements, you need to know how to handle significant figures. This is especially true when you are using a calculator to carry out mathematical operations. The answers given on a calculator can be derived results with more digits than are justified by the measurements.

Suppose you used a calculator to divide a measured value of 154 g by a measured value of 327 mL. Each of these values has three significant figures. The calculator would show a numerical answer of 0.470948012. The answer contains digits not justified by the measurements used to calculate it. Such an answer has to be rounded off to make its degree of certainty match that in the original measurements. The answer should be 0.471 g/mL.

The rules for rounding are shown in **Table 6.** The extent of rounding required in a given case depends on whether the numbers are being added, subtracted, multiplied, or divided.

TABLE 6 ***Rules for Rounding Numbers***

If the digit following the last digit to be retained is:	then the last digit should:	Example (rounded to three significant figures)
greater than 5	be increased by 1	42.68 g $\longrightarrow$ 42.7 g
less than 5	stay the same	17.32 m $\longrightarrow$ 17.3 m
5, followed by nonzero digit(s)	be increased by 1	2.7851 cm $\longrightarrow$ 2.79 cm
5, not followed by nonzero digit(s), and preceded by an odd digit	be increased by 1	4.635 kg $\longrightarrow$ 4.64 kg (because 3 is odd)
5, not followed by nonzero digit(s), and the preceding significant digit is even	stay the same	78.65 mL $\longrightarrow$ 78.6 mL (because 6 is even)

Addition or Subtraction with Significant Figures

Consider two mass measurements, 25.1 g and 2.03 g. The first measurement, 25.1 g, has one digit to the right of the decimal point, in the tenths place. There is no information on possible values for the hundredths place. That place is simply blank and cannot be assumed to be zero. The other measurement, 2.03 g, has two digits to the right of the decimal point. It provides information up to and including the hundredths place.

Suppose you were asked to add the two measurements. Simply carrying out the addition would result in an answer of 25.1 g + 2.03 g = 27.13 g. That answer suggests there is certainty all the way to the hundredths place. However, that result is not justified because the hundredths place in 25.1 g is completely unknown. The answer must be adjusted to reflect the uncertainty in the numbers added.

When adding or subtracting decimals, the answer must have the same number of digits to the right of the decimal point as there are in the measurement having the fewest digits to the right of the decimal point. Comparing the two values 25.1 g and 2.03 g, the measurement with the fewest digits to the right of the decimal point is 25.1 g. It has only one such digit. Following the rule, the answer must be rounded so that it has no more than one digit to the right of the decimal point. The answer should therefore be rounded to 27.1 g. When working with whole numbers, the answer should be rounded so that the final significant digit is in the same place as the leftmost uncertain digit. (For example, 5400 + 365 = 5800.)

Multiplication and Division with Significant Figures

Suppose you calculated the density of an object that has a mass of 3.05 g and a volume of 8.47 mL. The following division on a calculator will give a value of 0.360094451.

$$density = \frac{mass}{volume} = \frac{3.05 \text{ g}}{8.47 \text{ mL}} = 0.360094451 \text{ g/mL}$$

The answer must be rounded to the correct number of significant figures. The values of mass and volume used to obtain the answer have only three significant figures each. The degree of certainty in the calculated result is not justified. *For multiplication or division, the answer can have no more significant figures than are in the measurement with the fewest number of significant figures*. In the calculation just described, the answer, 0.360094451 g/mL, would be rounded to three significant figures to match the significant figures in 8.47 mL and 3.05 g. The answer would thus be 0.360 g/mL.

SAMPLE PROBLEM E *For more help, go to the Math Tutor at the end of Chapter 1.*

Carry out the following calculations. Express each answer to the correct number of significant figures.

a. 5.44 m − 2.6103 m

b. 2.4 g/mL × 15.82 mL

SOLUTION Carry out each mathematical operation. Follow the rules in **Table 5** and **Table 6** for determining significant figures and for rounding.

a. The answer is rounded to 2.83 m, because for subtraction there should be two digits to the right of the decimal point, to match 5.44 m.

b. The answer is rounded to 38 g, because for multiplication there should be two significant figures in the answer, to match 2.4 g/mL.

PRACTICE *Answers in Appendix E*

1. What is the sum of 2.099 g and 0.05681 g?
2. Calculate the quantity 87.3 cm − 1.655 cm.
3. Calculate the area of a rectangular crystal surface that measures 1.34 µm by 0.7488 µm. (Hint: Recall that *area* = *length* × *width* and is measured in square units.)
4. Polycarbonate plastic has a density of 1.2 g/cm^3. A photo frame is constructed from two 3.0 mm sheets of polycarbonate. Each sheet measures 28 cm by 22 cm. What is the mass of the photo frame?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate using significant figures.

Keyword: HC6MEAX

Conversion Factors and Significant Figures

Earlier in this chapter, you learned how conversion factors are used to change one unit to another. Such conversion factors are typically exact. That is, there is no uncertainty in them. For example, there are exactly 100 cm in a meter. If you were to use the conversion factor 100 cm/m to change meters to centimeters, the 100 would not limit the degree of certainty in the answer. Thus, 4.608 m could be converted to centimeters as follows.

$$4.608 \text{ m} \times \frac{100 \text{ cm}}{\text{m}} = 460.8 \text{ cm}$$

The answer still has four significant figures. Because the conversion factor is considered exact, the answer would not be rounded. Most exact conversion factors are defined, rather than measured, quantities. Counted numbers also produce conversion factors of unlimited precision. For example, if you counted that there are 10 test tubes for every student, that would produce an exact conversion factor of 10 test tubes/student. There is no uncertainty in that factor.

Scientific Notation

In **scientific notation,** *numbers are written in the form* M × *10*^{n}*, where the factor* M *is a number greater than or equal to 1 but less than 10 and* n *is a whole number.* For example, to write the quantity 65 000 km in scientific notation and show the first two digits as significant, you would write the following.

$$6.5 \times 10^4 \text{ km}$$

Writing the M factor as 6.5 shows that there are exactly two significant figures. If, instead, you intended the first three digits in 65 000 to be significant, you would write 6.50×10^4 km. When numbers are written in scientific notation, only the significant figures are shown.

Suppose you are expressing a very small quantity, such as the length of a flu virus. In ordinary notation this length could be 0.000 12 mm. That length can be expressed in scientific notation as follows.

$$0.000\,12 \text{ mm} = 1.2 \times 10^{-4} \text{ mm}$$

Move the decimal point four places to the right, and multiply the number by 10^{-4}.

1. Determine M by moving the decimal point in the original number to the left or the right so that only one nonzero digit remains to the left of the decimal point.
2. Determine n by counting the number of places that you moved the decimal point. If you moved it to the left, n is positive. If you moved it to the right, n is negative.

Mathematical Operations Using Scientific Notation

1. *Addition and subtraction* These operations can be performed only if the values have the same exponent (n factor). If they do not, adjustments must be made to the values so that the exponents are equal. Once the exponents are equal, the M factors can be added or subtracted. The exponent of the answer can remain the same, or it may then require adjustment if the M factor of the answer has more than one digit to the left of the decimal point. Consider the example of the addition of 4.2×10^4 kg and 7.9×10^3 kg.

   We can make both exponents either 3 or 4. The following solutions are possible.

$$\begin{array}{r} 4.2 \times 10^4 \text{ kg} \\ +0.79 \times 10^4 \text{ kg} \\ \hline 4.99 \times 10^4 \text{ kg} \end{array} \text{ rounded to } 5.0 \times 10^4 \text{ kg}$$

or

$$\begin{array}{r} 7.9 \times 10^3 \text{ kg} \\ +42 \times 10^3 \text{ kg} \\ \hline 49.9 \times 10^3 \text{ kg} \end{array} = 4.99 \times 10^4 \text{ kg rounded to } 5.0 \times 10^4 \text{ kg}$$

   Note that the units remain kg throughout.

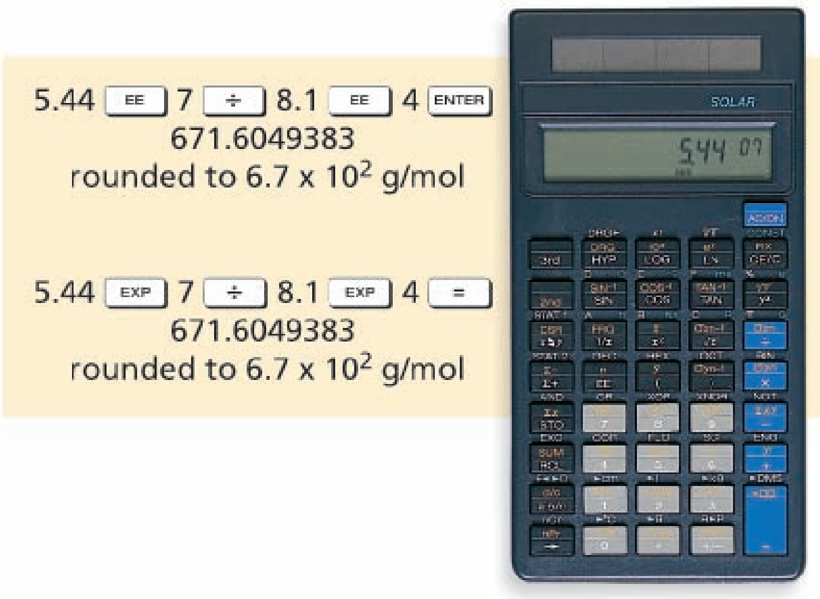


FIGURE 10 When you use a scientific calculator to work problems in scientific notation, don't forget to express the value on the display to the correct number of significant figures and show the units when you write the final answer.

2. *Multiplication* The *M* factors are multiplied, and the exponents are added algebraically.

Consider the multiplication of 5.23×10^6 µm by 7.1×10^{-2} µm.

$$\begin{aligned}(5.23 \times 10^6\ \mu\text{m})(7.1 \times 10^{-2}\ \mu\text{m}) &= (5.23 \times 7.1)(10^6 \times 10^{-2})\\ &= 37.133 \times 10^4\ \mu\text{m}^2 \text{ (adjust to two significant digits)}\\ &= 3.7 \times 10^5\ \mu\text{m}^2\end{aligned}$$

Note that when length measurements are multiplied, the result is area. The unit is now µm^2.

3. *Division* The *M* factors are divided, and the exponent of the denominator is subtracted from that of the numerator. The calculator keystrokes for this problem are shown in **Figure 10.**

$$\begin{aligned}\frac{5.44 \times 10^7\ \text{g}}{8.1 \times 10^4\ \text{mol}} &= \frac{5.44}{8.1} \times 10^{7-4}\ \text{g/mol}\\ &= 0.6716049383 \times 10^3 \text{ (adjust to two significant figures)}\\ &= 6.7 \times 10^2\ \text{g/mol}\end{aligned}$$

Note that the unit for the answer is the ratio of grams to moles.

Using Sample Problems

Learning to analyze and solve such problems requires practice and a logical approach. In this section, you will review a process that can help you analyze problems effectively. Most Sample Problems in this book are organized by four basic steps to guide your thinking in how to work out the solution to a problem.

Analyze

The first step in solving a quantitative word problem is to read the problem carefully at least twice and to analyze the information in it. Note any important descriptive terms that clarify or add meaning to the problem. Identify and list the data given in the problem. Also identify the unknown—the quantity you are asked to find.

Plan

The second step is to develop a plan for solving the problem. The plan should show how the information given is to be used to find the unknown. In the process, reread the problem to make sure you have gathered all the necessary information. It is often helpful to draw a picture that represents the problem. For example, if you were asked to determine the volume of a crystal given its dimensions, you could draw a representation of the crystal and label the dimensions. This drawing would help you visualize the problem.

Decide which conversion factors, mathematical formulas, or chemical principles you will need to solve the problem. Your plan might suggest a single calculation or a series of them involving different conversion factors. Once you understand how you need to proceed, you may wish to sketch out the route you will take, using arrows to point the way from one stage of the solution to the next. Sometimes you will need data that are not actually part of the problem statement. For instance, you'll often use data from the periodic table.

Compute

The third step involves substituting the data and necessary conversion factors into the plan you have developed. At this stage you calculate the answer, cancel units, and round the result to the correct number of significant figures. It is very important to have a plan worked out in step 2 before you start using the calculator. All too often, students start multiplying or dividing values given in the problem before they really understand what they need to do to get an answer.

Evaluate

Examine your answer to determine whether it is reasonable. Use the following methods, when appropriate, to carry out the evaluation.

1. Check to see that the units are correct. If they are not, look over the setup. Are the conversion factors correct?
2. Make an estimate of the expected answer. Use simpler, rounded numbers to do so. Compare the estimate with your actual result. The two should be similar.
3. Check the order of magnitude in your answer. Does it seem reasonable compared with the values given in the problem? If you calculated the density of vegetable oil and got a value of 54.9 g/mL, you should know that something is wrong. Oil floats on water; therefore, its density is less than water, so the value obtained should be less than 1.0 g/mL.
4. Be sure that the answer given for any problem is expressed using the correct number of significant figures.

Look over the following quantitative Sample Problem. Notice how the four-step approach is used, and then apply the approach yourself in solving the practice problems that follow.

SAMPLE PROBLEM F

Calculate the volume of a sample of aluminum that has a mass of 3.057 kg. The density of aluminum is 2.70 g/cm³.

SOLUTION

1 ANALYZE

Given: mass = 3.057 kg, density = 2.70 g/cm^3

Unknown: volume of aluminum

2 PLAN

The density unit in the problem is g/cm^3, and the mass given in the problem is expressed in kg. Therefore, in addition to using the density equation, you will need a conversion factor representing the relationship between grams and kilograms.

$$1000 \text{ g} = 1 \text{ kg}$$

Also, rearrange the density equation to solve for volume.

$$density = \frac{mass}{volume} \quad \text{or} \quad D = \frac{m}{V}$$

$$V = \frac{m}{D}$$

3 COMPUTE

$$V = \frac{3.057 \cancel{\text{kg}}}{2.70 \cancel{\text{g}}/\text{cm}^3} \times \frac{1000 \cancel{\text{g}}}{\cancel{\text{kg}}} = 1132.222 \ldots \text{cm}^3 \text{ (calculator answer)}$$

The answer should be rounded to three significant figures.

$$V = 1.13 \times 10^3 \text{ cm}^3$$

4 EVALUATE

The unit of volume, cm^3, is correct. An order-of-magnitude estimate would put the answer at over 1000 cm^3.

$$\frac{3}{2} \times 1000$$

The correct number of significant figures is three, which matches that in 2.70 g/cm^3.

PRACTICE

Answers in Appendix E

1. What is the volume, in milliliters, of a sample of helium that has a mass of 1.73×10^{-3} g, given that the density is 0.178 47 g/L?
2. What is the density of a piece of metal that has a mass of 6.25×10^5 g and is 92.5 cm × 47.3 cm × 85.4 cm?
3. How many millimeters are there in 5.12×10^5 kilometers?
4. A clock gains 0.020 second per minute. How many seconds will the clock gain in exactly six months, assuming exactly 30 days per month?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate using scientific notation.

Keyword: HC6MEAX

Direct Proportions

Two quantities are **directly proportional** *to each other if dividing one by the other gives a constant value*. For example, if the masses and volumes of different samples of aluminum are measured, the masses and volumes will be directly proportional to each other. As the masses of the samples increase, their volumes increase by the same factor, as you can see from the data in **Table 7.** Doubling the mass doubles the volume. Halving the mass halves the volume.

When two variables, x and y, are directly proportional to each other, the relationship can be expressed as $y \propto x$, which is read as "y is *proportional* to x." The general equation for a directly proportional relationship between the two variables can also be written as follows.

$$\frac{y}{x} = k$$

The value of k is a constant called the proportionality constant. Written in this form, the equation expresses an important fact about direct proportion: the ratio between the variables remains constant. Note that using the mass and volume values in **Table 7** gives a mass-volume ratio that is constant (neglecting measurement error). The equation can be rearranged into the following form.

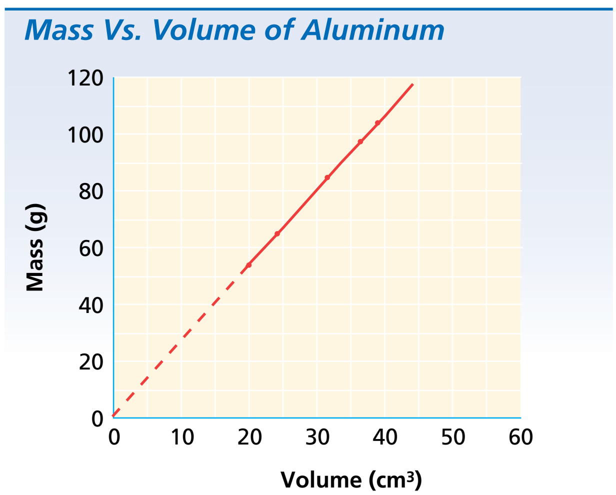
$$y = kx$$

The equation $y = kx$ may look familiar to you. It is the equation for a special case of a straight line. If two variables related in this way are graphed versus one another, a straight line, or linear plot that passes through the origin (0,0), results. The data for aluminum from **Table 7** are graphed in **Figure 11.** The mass and volume of a pure substance are directly proportional to each other. Consider mass to be y and volume to be x. The constant ratio, k, for the two variables is density. The slope of the line reflects the constant density, or mass-volume ratio, of aluminum,

TABLE 7 Mass-Volume Data for Aluminum at 20°C

Mass (g)	Volume (cm^3)	$\frac{m}{V}$ (g/cm^3)
54.7	20.1	2.72
65.7	24.4	2.69
83.5	30.9	2.70
96.3	35.8	2.69
105.7	39.1	2.70

FIGURE 11 The graph of mass versus volume shows a relationship of direct proportion. Notice that the line is extrapolated to pass through the origin.



which is 2.70 g/cm^3 at 20°C. Notice also that the plotted line passes through the origin. All directly proportional relationships produce linear graphs that pass through the origin.

Inverse Proportions

Two quantities are **inversely proportional** *to each other if their product is constant*. An example of an inversely proportional relationship is that between speed of travel and the time required to cover a fixed distance. The greater the speed, the less time that is needed to go a certain fixed distance. Doubling the speed cuts the required time in half. Halving the speed doubles the required time.

When two variables, x and y, are inversely proportional to each other, the relationship can be expressed as follows.

$$y \propto \frac{1}{x}$$

This is read "y is *proportional* to 1 divided by x." The general equation for an inversely proportional relationship between the two variables can be written in the following form.

$$xy = k$$

In the equation, k is the proportionality constant. If x increases, y must decrease by the same factor to keep the product constant.

A graph of variables that are inversely proportional produces a curve called a hyperbola. Such a graph is illustrated in **Figure 12.** When the temperature of the gas is kept constant, the volume (V) of the gas sample decreases as the pressure (P) increases. Look at the data shown in **Table 8.** Note that $P \times V$ gives a reasonably constant value. The graph of this data is shown in **Figure 12.**

TABLE 8 Pressure-Volume Data for Nitrogen at Constant Temperature

Pressure (kPa)	Volume (cm^3)	$P \times V$
100	500	50 000
150	333	50 000
200	250	50 000
250	200	50 000
300	166	49 800
350	143	50 100
400	125	50 000
450	110	49 500

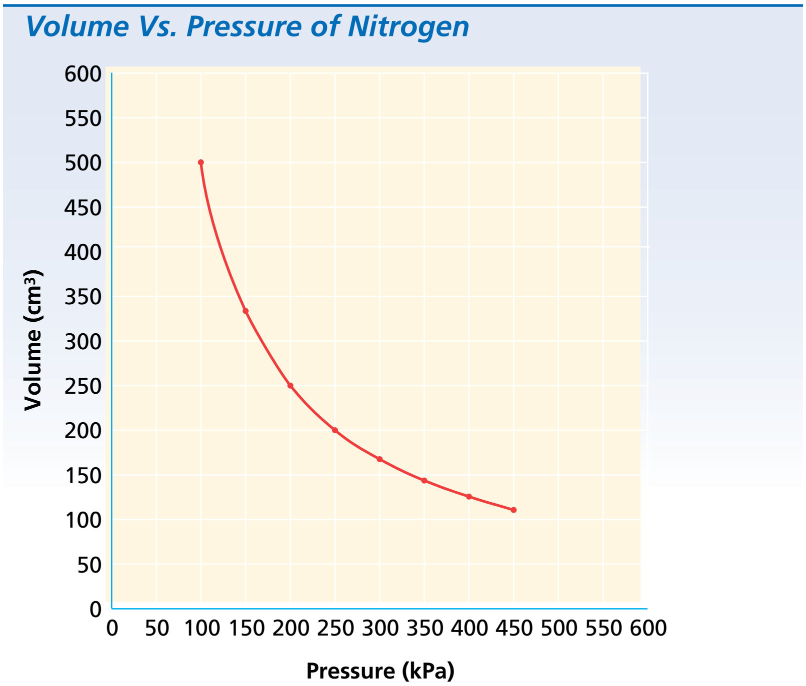


FIGURE 12 The graph of volume versus pressure shows an inversely proportional relationship. Note the difference between the shape of this graph and that of the graph in Figure 11.

SECTION REVIEW

1. The density of copper is listed as 8.94 g/cm^3. Two students each make three density determinations of samples of the substance. Student A's results are 7.3 g/mL, 9.4 g/mL, and 8.3 g/mL. Student B's results are 8.4 g/cm^3, 8.8 g/cm^3, and 8.0 g/cm^3. Compare the two sets of results in terms of precision and accuracy.
2. Determine the number of significant figures.
   a. 6.002 cm
   b. 0.0020 m
   c. 10.0500 g
   d. 7000 kg
   e. 7000. kg
3. Round 2.6765 to two significant figures.
4. Carry out the following calculations.
   a. 52.13 g + 1.7502 g
   b. 12 m $\times$ 6.41 m
   c. $\dfrac{16.25 \text{ g}}{5.1442 \text{ mL}}$
5. Perform the following operations. Express each answer in scientific notation.
   a. $(1.54 \times 10^{-2} \text{ g}) + (2.86 \times 10^{-1} \text{ g})$
   b. $(7.023 \times 10^{9} \text{ g}) - (6.62 \times 10^{7} \text{ g})$
   c. $(8.99 \times 10^{-4} \text{ m}) \times (3.57 \times 10^{4} \text{ m})$
   d. $\dfrac{2.17 \times 10^{-3} \text{ g}}{5.022 \times 10^{4} \text{ mL}}$
6. Write the following numbers in scientific notation.
   a. 560 000
   b. 33 400
   c. 0.000 4120
7. A student measures the mass of a beaker filled with corn oil. The mass reading averages 215.6 g. The mass of the beaker is 110.4 g.
   a. What is the mass of the corn oil?
   b. What is the density of the corn oil if its volume is 114 cm^3?
8. Calculate the mass of gold that occupies 5.0×10^{-3} cm^3. The density of gold is 19.3 g/cm^3.
9. What is the difference between a graph representing data that are directly proportional and a graph of data that are inversely proportional?

Critical Thinking

10. **APPLYING CONCEPTS** The mass of a liquid is 11.50 g and its volume is 9.03 mL. How many significant figures should its density value have? Explain the reason for your answer.

CHAPTER HIGHLIGHTS

Scientific Method

Vocabulary
scientific method
system
hypothesis
model
theory

- The scientific method is a logical approach to solving problems that lend themselves to investigation.
- A hypothesis is a testable statement that serves as the basis for predictions and further experiments.
- A theory is a broad generalization that explains a body of known facts or phenomena.

Units of Measurement

Vocabulary
quantity
SI
weight
derived unit
volume
density
conversion factor
dimensional analysis

- The result of nearly every measurement is a number and a unit.
- The SI system of measurement is used in science. It has seven base units: the meter (length), kilogram (mass), second (time), kelvin (temperature), mole (amount of substance), ampere (electric current), and candela (luminous intensity).
- Weight is a measure of the gravitational pull on matter.
- Derived SI units include the square meter (area) and the cubic meter (volume).
- Density is the ratio of mass to volume.
- Conversion factors are used to convert from one unit to another.

Using Scientific Measurements

Vocabulary
accuracy
precision
percentage error
significant figures
scientific notation
directly proportional
inversely proportional

- Accuracy refers to the closeness of a measurement to the correct or accepted value. Precision refers to the closeness of values for a set of measurements.
- Percentage error is the difference between the experimental and the accepted value that is divided by the accepted value and then multiplied by 100.
- The significant figures in a number consist of all digits known with certainty plus one final digit, which is uncertain.
- After addition or subtraction, the answer should be rounded so that it has no more digits to the right of the decimal point than there are in the measurement that has the smallest number of digits to the right of the decimal point. After multiplication or division, the answer should be rounded so that it has no more significant figures than there are in the measurement that has the fewest number of significant figures.
- Exact conversion factors are completely certain and do not limit the number of digits in a calculation.
- A number written in scientific notation is of the form $M \times 10^n$, in which M is greater than or equal to 1 but less than 10 and n is an integer.
- Two quantities are directly proportional to each other if dividing one by the other yields a constant value. Two quantities are inversely proportional to each other if their product has a constant value.

CHAPTER REVIEW

For more practice, go to the Problem Bank in Appendix D.

Scientific Method

SECTION 1 REVIEW

1. How does quantitative information differ from qualitative information?
2. What is a hypothesis?
3. a. What is a model in the scientific sense?
   b. How does a model differ from a theory?

Units of Measurement

SECTION 2 REVIEW

4. Why is it important for a measurement system to have an international standard?
5. How does a quantity differ from a unit? Use two examples to explain the difference.
6. List the seven SI base units and the quantities they represent.
7. What is the numerical equivalent of each of the following SI prefixes?
   a. kilo-
   b. centi-
   c. mega-
   d. micro-
   e. milli-
8. Identify the SI unit that would be most appropriate for expressing the length of the following.
   a. width of a gymnasium
   b. length of a finger
   c. distance between your town and the closest border of the next state
   d. length of a bacterial cell
9. Identify the SI unit that would be most appropriate for measuring the mass of each of the following objects.
   a. table
   b. coin
   c. a 250 mL beaker
10. Explain why the second is not defined by the length of the day.
11. a. What is a derived unit?
    b. What is the SI-derived unit for area?
12. a. List two SI-derived units for volume.
    b. List two non-SI units for volume, and explain how they relate to the cubic centimeter.
13. a. Why are the units used that are to express the densities of gases different from those used to express the densities of solids or liquids?
    b. Name two units for density.
    c. Why is the temperature at which density is measured usually specified?
14. a. Which of the solids listed in **Table 4** will float on water?
    b. Which of the liquids will sink in milk?
15. a. Define *conversion factor.*
    b. Explain how conversion factors are used.

PRACTICE PROBLEMS

16. What is the volume, in cubic meters, of a rectangular solid that is 0.25 m long, 6.1 m wide, and 4.9 m high?
17. Find the density of a material, given that a 5.03 g sample occupies 3.24 mL. (Hint: See Sample Problem A.)
18. What is the mass of a sample of material that has a volume of 55.1 cm^3 and a density of 6.72 g/cm^3?
19. A sample of a substance that has a density of 0.824 g/mL has a mass of 0.451 g. Calculate the volume of the sample.
20. How many grams are in 882 μg? (Hint: See Sample Problem B.)
21. Calculate the number of milliliters in 0.603 L.
22. The density of gold is 19.3 g/cm^3.
    a. What is the volume, in cubic centimeters, of a sample of gold that has a mass of 0.715 kg?
    b. If this sample of gold is a cube, what is the length of each edge in centimeters?
23. a. Find the number of kilometers in 92.25 m.
    b. Convert the answer in kilometers to centimeters.

Using Scientific Measurements

SECTION 3 REVIEW

24. Compare accuracy and precision.
25. a. Write the equation that is used to calculate percentage error.

b. Under what condition will percentage error be negative?

26. How is the average for a set of values calculated?

27. What is meant by a mass measurement expressed in this form: 4.6 g ± 0.2 g?

28. Suppose a graduated cylinder were not correctly calibrated. How would this affect the results of a measurement? How would it affect the results of a calculation using this measurement?

29. Round each of the following measurements to the number of significant figures indicated.
a. 67.029 g to three significant figures
b. 0.15 L to one significant figure
c. 52.8005 mg to five significant figures
d. 3.174 97 mol to three significant figures

30. State the rules governing the number of significant figures that result from each of the following operations.
a. addition and subtraction
b. multiplication and division

31. What is the general form for writing numbers in scientific notation?

32. a. By using x and y, state the general equation for quantities that are directly proportional.
b. For two directly proportional quantities, what happens to one variable when the other variable increases?

33. a. State the general equation for quantities, x and y, that are inversely proportional.
b. For two inversely proportional quantities, what happens to one variable when the other increases?

34. Arrange in the correct order the following four basic steps for finding the solution to a problem: compute, plan, evaluate, and analyze.

PRACTICE PROBLEMS

35. A student measures the mass of a sample as 9.67 g. Calculate the percentage error, given that the correct mass is 9.82 g. (Hint: See Sample Problem C.)

36. A handbook gives the density of calcium as 1.54 g/cm^3. Based on lab measurements, what is the percentage error of a density calculation of 1.25 g/cm^3?

37. What is the percentage error of a length measurement of 0.229 cm if the correct value is 0.225 cm?

38. How many significant figures are in each of the following measurements? (Hint: See Sample Problem D.)
a. 0.4004 mL
b. 6000 g
c. 1.000 30 km
d. 400. mm

39. Calculate the sum of 6.078 g and 0.3329 g.

40. Subtract 7.11 cm from 8.2 cm. (Hint: See Sample Problem E.)

41. What is the product of 0.8102 m and 3.44 m?

42. Divide 94.20 g by 3.167 22 mL.

43. Write the following numbers in scientific notation.
a. 0.000 673 0
b. 50 000.0
c. 0.000 003 010

44. The following numbers are in scientific notation. Write them in ordinary notation.
a. 7.050×10^3 g
b. $4.000\ 05 \times 10^7$ mg
c. $2.350\ 0 \times 10^4$ mL

45. Perform the following operation. Express the answer in scientific notation and with the correct number of significant figures.
$0.002115\text{m} \times 0.0000405\text{m}$

46. A sample of a certain material has a mass of 2.03×10^{-3} g. Calculate the volume of the sample, given that the density is 9.133×10^{-1} g/cm^3. Use the four-step method to solve the problem. (Hint: See Sample Problem F.)

MIXED REVIEW

47. A man finds that he has a mass of 100.6 kg. He goes on a diet, and several months later he finds that he has a mass of 96.4 kg. Express each number in scientific notation, and calculate the number of kilograms the man has lost by dieting.

48. A large office building is 1.07×10^2 m long, 31 m wide, and 4.25×10^2 m high. What is its volume?

49. An object has a mass of 57.6 g. Find the object's density, given that its volume is 40.25 cm^3.

50. A lab worker measures the mass of some sucrose as 0.947 mg. Convert that quantity to grams and to kilograms.

51. A student calculates the density of iron as 6.80 g/cm^3 by using lab data for mass and volume. A handbook reveals that the correct value is 7.86 g/cm^3. What is the percentage error?

USING THE HANDBOOK

52. Find the table of properties for Group 1 elements in the *Elements Handbook*. Calculate the volume of a single atom of each element listed in the table by using the equation for the volume of a sphere.

$$\frac{4}{3}\pi \cdot r^3$$

53. Use the radius of a sodium atom from the *Elements Handbook* to calculate the number of sodium atoms in a row 5.00 cm long. Assume that each sodium atom touches the ones next to it.

54. a. A block of sodium that has the measurements 3.00 cm × 5.00 cm × 5.00 cm has a mass of 75.5 g. Calculate the density of sodium.

b. Compare your calculated density with the value in the properties table for Group 1 elements. Calculate the percentage error for your density determination.

RESEARCH & WRITING

55. How does the metric system, which was once a standard for measurement, differ from SI? Why was it necessary for the United States to change to SI?

56. What are ISO 9000 standards? How do they affect industry on an international level?

ALTERNATIVE ASSESSMENT

57. Performance Obtain three metal samples from your teacher. Determine the mass and volume of each sample. Calculate the density of each metal from your measurement data. (Hint: Consider using the water displacement technique to measure the volume of your samples.)

58. Use the data from the Nutrition Facts label below to answer the following questions:

a. Use the data given on the label for grams of fat and Calories from fat to construct a conversion factor that has the units Calories per gram.

b. Calculate the mass in kilograms for 20 servings of the food.

c. Calculate the mass of protein in micrograms for one serving of the food.

d. What is the correct number of significant figures for the answer in item a? Why?

Nutrition Facts

Serving Size ¾ cup (30g)

Servings Per Container About 14

Amount Per Serving	Corn Crunch	with ½ cup skim milk
Calories	120	160
Calories from Fat	15	20
	% **Daily Value****	
Total Fat 2g*	**3**%	**3**%
Saturated Fat 0g	**0**%	**0**%
Cholesterol 0mg	**0**%	**1**%
Sodium 160mg	**7**%	**9**%
Potassium 65mg	**2**%	**8**%
Total Carbohydrate 25g	**8**%	**10**%
Dietary Fiber 3g		
Sugars 3g		
Other Carbohydrate 11g		
Protein 2g		

*Amount in Cereal. A serving of cereal plus skim milk provides 2g fat, less 5mg cholesterol, 220mg sodium, 270mg potassium, 31g carbohydrate (19g sugars) and 6g protein.

**Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs:

	Calories	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	25g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Potassium		3,500mg	3,500mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

Math Tutor SCIENTIFIC NOTATION

Any value expressed in scientific notation, whether large or small, has two parts. The first part, the *first factor,* consists of a number greater than or equal to 1 but less than 10, which may have any number of digits after the decimal point. The second part consists of a power of 10.

$$\underbrace{6.02}_{\text{first factor}} \times \underbrace{10^{\overbrace{23}^{\text{exponent}}}}_{\text{power of ten}}$$

To write the first part, move the decimal point to the right or the left so that there is only one nonzero digit to the left of the decimal point. The second part is written as an exponent, which is determined by counting the number of places the decimal point must be moved. If it is moved to the right, the exponent is negative. If it is moved to the left, the exponent is positive.

Problem-Solving *TIPS*

- In addition and subtraction, all values must first be converted to numbers that have the same exponent of 10. The result is the sum or the difference of the first factors, multiplied by the same exponent of 10. Finally, the result should be rounded to the correct number of significant figures and expressed in scientific notation.
- In multiplication, the first factors are multiplied and the exponents of 10 are added.
- In division, the *first factors* of the numbers are divided and the exponent of 10 in the denominator is subtracted from the exponent of 10 in the numerator.

SAMPLE 1

Write 299 800 000 m/s in scientific notation.

The decimal must move to the left 8 places, which indicates a positive exponent.

$$299\,800\,000.\ \text{m/s}$$

8 7 6 5 4 3 2 1

The value in scientific notation is 2.998×10^8 m/s.

SAMPLE 2

Solve the following equation and write the answer in scientific notation.

$$(3.1 \times 10^3)(5.21 \times 10^4)$$

Multiply the first factors, and then add the exponents of 10.

$$(3.1 \times 5.21) \times 10^{(3+4)} = 16 \times 10^7 = 1.6 \times 10^8$$

PRACTICE PROBLEMS

1. Rewrite the following numbers in scientific notation.

a. 0.0000745 g

b. 5984102 nm

2. Solve the following equations, and write the answers in scientific notation.

a. $1.017 \times 10^3 - 1.013 \times 10^4$

b. $\dfrac{9.27 \times 10^4}{11.24 \times 10^5}$

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Which of the following masses is the largest?
A. 0.200 g
B. 0.020 kg
C. 20.0 mg
D. 2000 µg

2. Which of the following measurements contains three significant figures?
A. 200 mL
B. 0.02 mL
C. 20.2 mL
D. 200.0 mL

3. A theory differs from a hypothesis in that a theory
A. cannot be disproved.
B. always leads to the formation of a law.
C. has been subjected to experimental testing.
D. represents an educated guess.

4. All measurements in science
A. must be expressed in scientific notation.
B. have some degree of uncertainty.
C. are both accurate and precise.
D. must include only those digits that are known with certainty.

5. When numbers are multiplied or divided, the answer can have no more
A. significant figures than are in the measurement that has the smallest number of significant figures.
B. significant figures than are in the measurement that has the largest number of significant figures.
C. digits to the right of the decimal point than are in the measurement that has the smallest number of digits to the right of the decimal point.
D. digits to the right of the decimal point than are in the measurement that has the largest number of digits to the right of the decimal point.

6. Which of the following is not part of the scientific method?
A. making measurements
B. introducing bias
C. making an educated guess
D. analyzing data

7. The accuracy of a measurement
A. is how close it is to the true value.
B. does not depend on the instrument used to measure the object.
C. indicates that the measurement is also precise.
D. is something that scientists rarely achieve.

8. A measurement of 23 465 mg converted to grams equals
A. 2.3465 g.
B. 23.465 g.
C. 234.65 g.
D. 0.23465 g.

9. A metal sample has a mass of 45.65 g. The volume of the sample is 16.9 cm^3. The density of the sample is
A. 2.7 g/cm^3.
B. 2.70 g/cm^3.
C. 0.370 g/cm^3.
D. 0.37 g/cm^3.

SHORT ANSWER

10. A recipe for 18 cookies calls for 1 cup of chocolate chips. How many cups of chocolate chips are needed for 3 dozen cookies? What kind of proportion, direct or indirect, did you use to answer this question?

11. Which of the following statements contain exact numbers?
A. There are 12 eggs in a dozen.
B. The accident injured 21 people.
C. The circumference of the Earth at the equator is 40 000 km.

EXTENDED RESPONSE

12. You have decided to test the effects of five garden fertilizers by applying some of each to separate rows of radishes. What is the variable you are testing? What factors should you control? How will you measure and analyze the results?

13. Around 1150, King David I of Scotland defined the inch as the width of a man's thumb at the base of the nail. Discuss the practical limitations of this early unit of measurement.

Test TIP Carefully read questions that ask for the one choice that is not correct. Often these questions include words such as *is not, except,* or *all but.*

Percentage of Water in Popcorn

OBJECTIVES

- ***Measure*** **the masses of various combinations of a beaker, oil, and popcorn kernels.**
- ***Determine*** **the percentages of water in popcorn kernels.**
- ***Compare*** **experimental data.**

MATERIALS

- **aluminum foil (1 sheet)**
- **beaker, 250 mL**
- **Bunsen burner with gas tubing and striker**
- **kernels of popcorn for each of three brands (80)**
- **oil to coat the bottom of the beaker**
- **ring stand, iron ring, and wire gauze**

BACKGROUND

Popcorn pops because of the natural moisture inside each kernel. When the internal water is heated above 100°C, the liquid water changes to a gas, which takes up much more space than the liquid, so the kernel expands rapidly.

The percentage of water in popcorn can be determined by the following equation.

$$\frac{\text{initial mass} - \text{final mass}}{\text{initial mass}} \times 100 = \text{percentage of } H_2O \text{ in unpopped popcorn}$$

The popping process works best when the kernels are first coated with a small amount of vegetable oil. Make sure you account for the presence of this oil when measuring masses. In this lab, you will design a procedure for determining the percentage of water in three samples of popcorn. The popcorn is for testing only, and *must not* be eaten.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. In your notebook, prepare a data table like the one your teacher has made on the board. The table should have three columns, one for each brand of popcorn.

PROCEDURE

1. Measure the mass of a 250 mL beaker. Record the mass in your data table.
2. Add a small amount of vegetable oil to the beaker to coat the bottom of it. Measure the mass of the beaker and oil. Record the mass in your data table.
3. Add 20 kernels of brand A popcorn to the beaker. Shake the beaker gently to coat the kernels with oil. Measure the mass of the beaker, oil, and popcorn. Record the mass in your data table.
4. Subtract the mass found in step 2 from the mass found in step 3 to obtain the mass of 20 unpopped kernels. Record the mass in your data table.
5. Cover the beaker loosely with the aluminum foil. Punch a few small holes in the aluminum foil to let moisture escape. These holes should not be large enough to let the popping corn pass through.
6. Heat the popcorn until the majority of the kernels have popped. The popcorn pops more efficiently if the beaker is held firmly with tongs and gently shaken side to side on the wire gauze.
7. Remove the aluminum foil from the beaker and allow the beaker to cool for 10 minutes. Then, measure the mass of the beaker, oil, and popped corn. Record the mass in your data table.
8. Subtract the mass in step 7 from the mass in step 3 to obtain the mass of water that escaped when the corn popped. Record the mass in your data table.
9. Calculate the percentage of water in the popcorn.
10. Dispose of the popcorn in the designated container. Remove the aluminum foil, and set it aside. Clean the beaker, and dry it well. Alternatively, if your teacher approves, use a different 250 mL beaker.
11. Repeat steps 1–10 for brand B popcorn.
12. Repeat steps 1–10 for brand C popcorn.

CLEANUP AND DISPOSAL

13. Dispose of popped popcorn and aluminum foil in containers as directed by your instructor. Do not eat the popcorn.
14. Clean beakers. Return beakers and other equipment to the proper place.
15. Clean all work surfaces and personal protective equipment as directed by your instructor.
16. Wash your hands thoroughly before leaving the laboratory.

ANALYSIS AND INTERPRETATION

1. **Applying Ideas:** Determine the mass of the 20 unpopped kernels of popcorn for each brand of popcorn.
2. **Applying Ideas:** Determine the mass of the 20 popped kernels of popcorn for each brand of popcorn.
3. **Applying Ideas:** Determine the mass of the water that was lost when the popcorn popped for each brand.

CONCLUSIONS

1. **Analyzing Data:** Determine the percentage by mass of water in each brand of popcorn.
2. **Inferring Relationships:** Do all brands of popcorn contain the same percentage water?

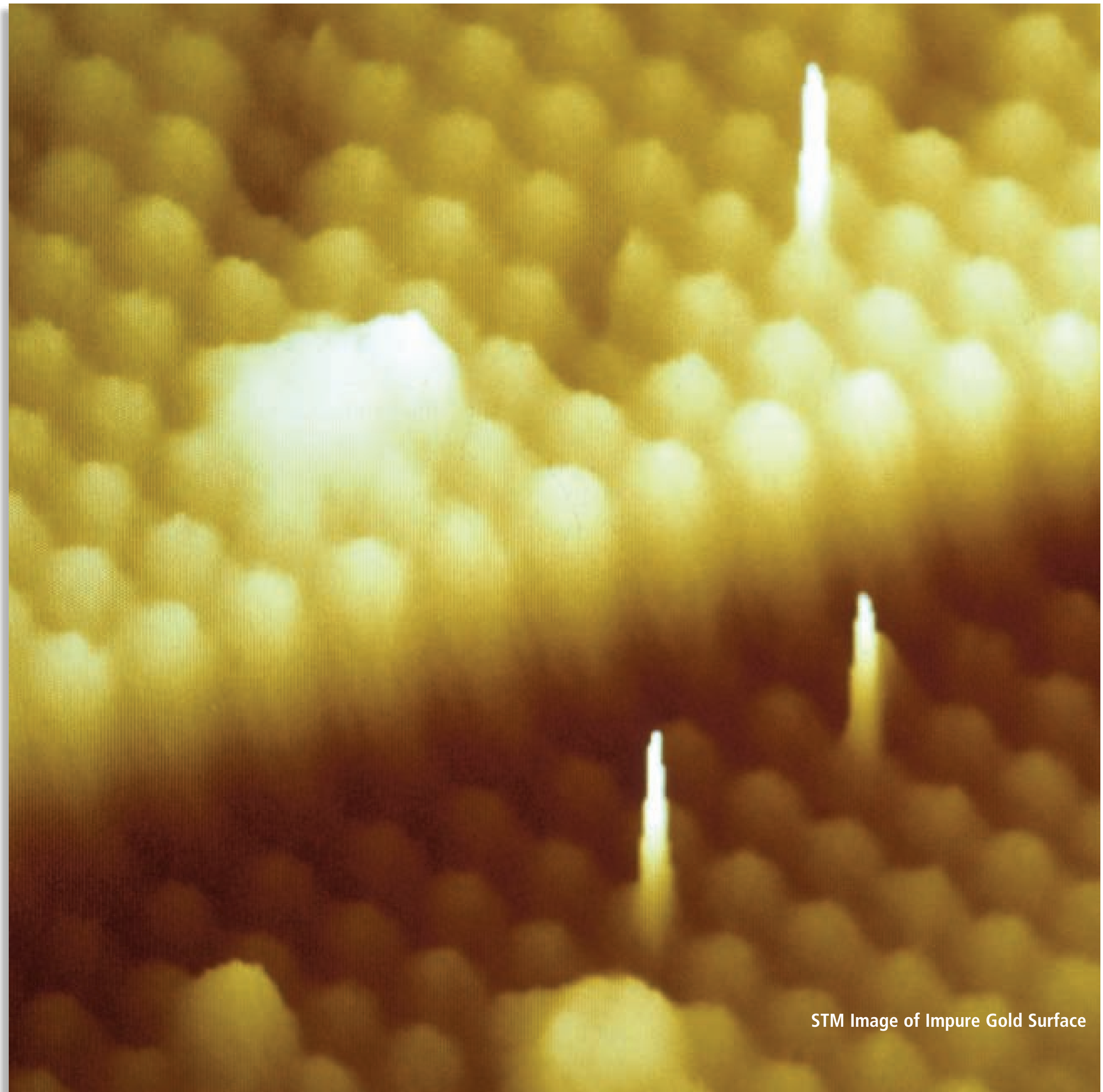
EXTENSIONS

1. **Designing Experiments:** What are some likely areas of imprecision in this experiment?
2. **Designing Experiments:** Do you think that the volume of popped corn depends on the percentage of water in the unpopped corn? Design an experiment to find the answer.

CHAPTER 3

Atoms: The Building Blocks of Matter

An atom is the smallest particle of an element that retains the chemical properties of that element.



STM Image of Impure Gold Surface

The Atom: From Philosophical Idea to Scientific Theory

SECTION 1

OBJECTIVES

- Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Summarize the five essential points of Dalton's atomic theory.
- Explain the relationship between Dalton's atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

When you crush a lump of sugar, you can see that it is made up of many smaller particles of sugar. You may grind these particles into a very fine powder, but each tiny piece is still sugar. Now suppose you dissolve the sugar in water. The tiny particles seem to disappear completely. Even if you look at the sugar-water solution through a powerful microscope, you cannot see any sugar particles. Yet if you were to taste the solution, you'd know that the sugar is still there. Observations like these led early philosophers to ponder the fundamental nature of matter. Is it continuous and infinitely divisible, or is it divisible only until a basic, invisible particle that cannot be divided further is reached?

The particle theory of matter was supported as early as 400 B.C. by certain Greek thinkers, such as Democritus. He called nature's basic particle an *atom,* based on the Greek word meaning "indivisible." Aristotle was part of the generation that succeeded Democritus. His ideas had a lasting impact on Western civilization, and he did not believe in atoms. He thought that all matter was continuous, and his opinion was accepted for nearly 2000 years. Neither the view of Aristotle nor that of Democritus was supported by experimental evidence, so each remained speculation until the eighteenth century. Then scientists began to gather evidence favoring the atomic theory of matter.

extension

Historical Chemistry

Go to **go.hrw.com** for a full-length article on the history of atomic theory and transmutation.

Keyword: HC6ATMX

Foundations of Atomic Theory

Virtually all chemists in the late 1700s accepted the modern definition of an element as a substance that cannot be further broken down by ordinary chemical means. It was also clear that elements combine to form compounds that have different physical and chemical properties than those of the elements that form them. There was great controversy, however, as to whether elements always combine in the same ratio when forming a particular compound.

The transformation of a substance or substances into one or more new substances is known as a *chemical reaction*. In the 1790s, the study of matter was revolutionized by a new emphasis on the quantitative

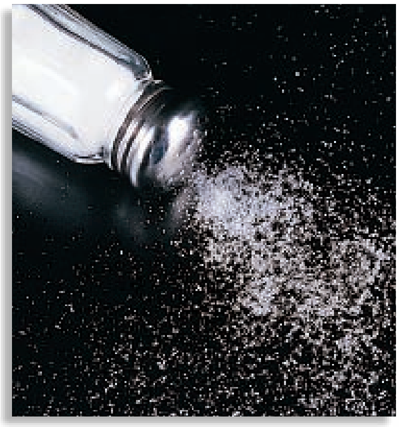


FIGURE 1 Each of the salt crystals shown here contains exactly 39.34% sodium and 60.66% chlorine by mass.

analysis of chemical reactions. Aided by improved balances, investigators began to accurately measure the masses of the elements and compounds they were studying. This led to the discovery of several basic laws. One of these laws was the **law of conservation of mass,** *which states that mass is neither created nor destroyed during ordinary chemical reactions or physical changes*. This discovery was soon followed by the assertion that, regardless of where or how a pure chemical compound is prepared, it is composed of a fixed proportion of elements. For example, sodium chloride, also known as ordinary table salt, *always* consists of 39.34% by mass of the element sodium, Na, and 60.66% by mass of the element chlorine, Cl. *The fact that a chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound is known as the* **law of definite proportions.**

It was also known that two elements sometimes combine to form more than one compound. For example, the elements carbon and oxygen form two compounds, carbon dioxide and carbon monoxide. Consider samples of each of these compounds, each containing 1.00 g of carbon. In carbon dioxide, 2.66 g of oxygen combine with 1.00 g of carbon. In carbon monoxide, 1.33 g of oxygen combine with 1.00 g of carbon. The ratio of the masses of oxygen in these two compounds is 2.66 to 1.33, or 2 to 1. This illustrates the **law of multiple proportions:** *If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers.*

Dalton's Atomic Theory

In 1808, an English schoolteacher named John Dalton proposed an explanation for the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. He reasoned that elements were composed of atoms and that only whole numbers of atoms can combine to form compounds. His theory can be summed up by the following statements.

1. All matter is composed of extremely small particles called atoms.
2. Atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged.

According to Dalton's atomic theory, the law of conservation of mass is explained by the fact that chemical reactions involve merely the combination, separation, or rearrangement of atoms and that during these processes atoms are not subdivided, created, or destroyed. This

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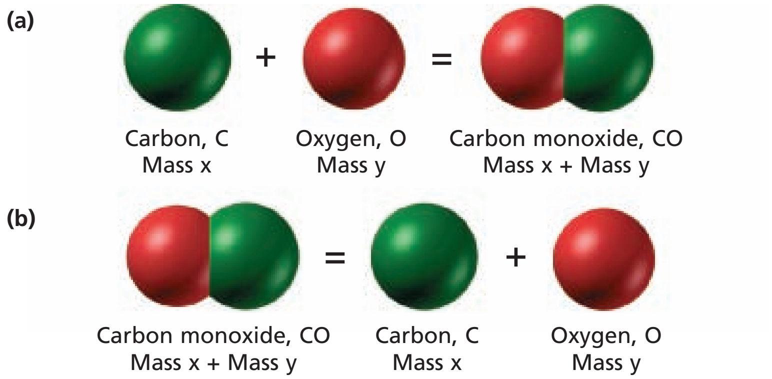


FIGURE 2 (a) An atom of carbon, C, and an atom of oxygen, O, can combine chemically to form a molecule of carbon monoxide, CO. The mass of the CO molecule is equal to the mass of the C atom plus the mass of the O atom. (b) The reverse holds true in a reaction in which a CO molecule is broken down into its elements.

idea is illustrated in **Figure 2** for the formation of carbon monoxide from carbon and oxygen.

The law of definite proportions, on the other hand, results from the fact that a given chemical compound is always composed of the same combination of atoms (see **Figure 3**). As for the law of multiple proportions, in the case of the carbon oxides, the 2-to-1 ratio of oxygen masses results because carbon dioxide always contains twice as many atoms of oxygen (per atom of carbon) as does carbon monoxide. This can also be seen in **Figure 3.**

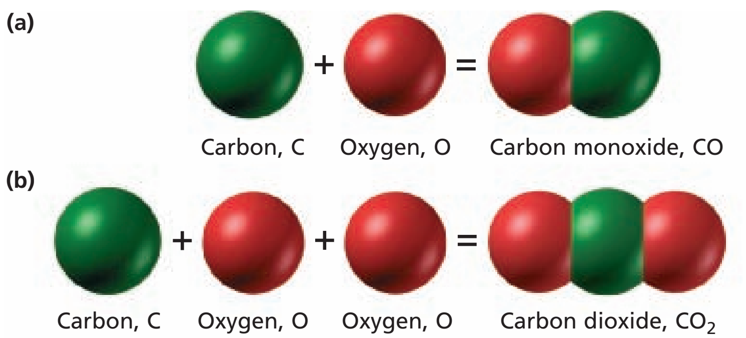
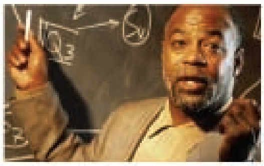


FIGURE 3 (a) CO molecules are always composed of one C atom and one O atom. (b) CO_2 molecules are always composed of one C atom and two O atoms. Note that a molecule of carbon dioxide contains twice as many oxygen atoms as does a molecule of carbon monoxide.

Modern Atomic Theory

By relating atoms to the measurable property of mass, Dalton turned Democritus's *idea* into a *scientific theory* that could be tested by experiment. But not all aspects of Dalton's atomic theory have proven to be correct. For example, today we know that atoms are divisible into even smaller particles (although the law of conservation of mass still holds true for chemical reactions). And, as you will see in Section 3, we know that a given element can have atoms with different masses. Atomic theory has not been discarded, however. Instead, it has been modified to explain the new observations. The important concepts that (1) all matter is composed of atoms and that (2) atoms of any one element differ in properties from atoms of another element remain unchanged.

CAREERS in Chemistry



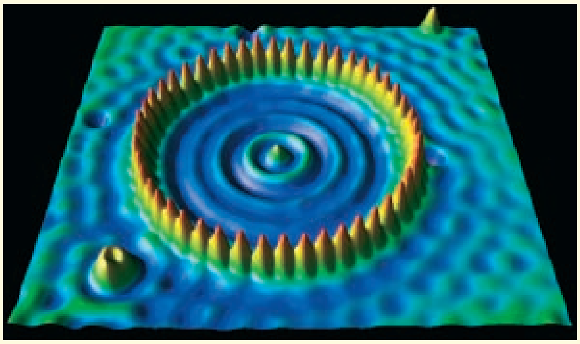
Physical Chemist

Physical chemists focus on understanding the physical properties of atoms and molecules. They are driven by a curiosity of what makes things work at the level of atoms, and they enjoy being challenged. In addition to chemistry, they study mathematics and physics extensively. Laboratory courses involving experience with electronics and optics are typically part of their training. Often, they enjoy working with instruments and computers. Physical chemists can be experimentalists or theoreticians. They use sophisticated instruments to make measurements, or high-powered computers to perform intensive calculations. The instruments used include lasers, electron microscopes, nuclear magnetic resonance spectrometers, mass spectrometers, and particle accelerators. Physical chemists work in industry, government laboratories, research institutes, and academic institutions. Because physical chemists work on a wide range of problems, taking courses in other science disciplines is important.

Scanning Tunneling Microscopy

For years, scientists have yearned for the ability to "see" individual atoms. Because atoms are so small, this had been nothing more than a dream. Now, the scanning tunneling microscope, STM, gives scientists the ability to look at individual atoms. It was invented in 1981 by Gerd Binnig and Heinrich Rohrer, scientists working for IBM in Zurich, Switzerland. They shared the 1986 Nobel Prize in physics for their discovery.

The basic principle of STM is based on the current that exists between a metallic needle that is sharpened to a single atom, the probe, and a conducting sample. As the probe passes above the surface of the sample at a distance of one or two atoms, electrons can "tunnel" from the needle tip to the sample's surface. The probe moves across, or "scans," the surface of the sample. When the probe comes close to the electrons of an individual atom, a signal is produced. A weaker signal is produced between atoms. These signals build a topographical (hill and valley) "map" of conducting and nonconducting regions. The resulting map shows the position and spacing of atoms.



▲ *This STM image shows a "corral" of iron atoms on a copper surface.*

Surface chemistry is a developing subdiscipline in physical chemistry, and STM is an important tool in the field. Scientists use STM to study surface reactions, such as those that take place in catalytic converters. Other areas of research in which STM is useful include semiconductors and microelectronics. Usually, STM is used with materials that conduct, but it has also been used to study biological molecules, such as DNA.

One innovative application of STM is the ability to position individual atoms. The figure shows the result of moving individual atoms. First, iron atoms were placed on a copper surface. Then, individual iron atoms were picked up by the probe and placed in position. The result is a "quantum corral" of 48 iron atoms on the surface of copper. The diameter of the corral is about 14 nm.

Questions

1. In addition to chemistry, what kinds of courses are important for a student interested in a physical chemistry career?
2. What part of an atom is detected by STM?

QuickLAB

Wear safety goggles and an apron.

Constructing a Model

Question

How can you construct a model of an unknown object by (1) making inferences about an object that is in a closed container and (2) touching the object without seeing it?

Procedure

Record all of your results in a data table.

1. Your teacher will provide you with a can that is covered by a sock sealed with tape. Without unsealing the container, try to determine the number of objects inside the can as well as the mass, shape, size, composition, and texture of each. To do this, you may carefully tilt or shake the can. Record your observations in a data table.
2. Remove the tape from the top of the sock. Do *not* look inside the can. Put one hand through the opening, and make the same observations as in step 1 by handling the objects. To make more-accurate estimations, practice estimating the sizes and masses of some known objects outside the can. Then compare your estimates of these objects with actual measurements using a metric ruler and a balance.

Discussion

1. Scientists often use more than one method to gather data. How was this illustrated in the investigation?
2. Of the observations you made, which were qualitative and which were quantitative?
3. Using the data you gathered, draw a model of the unknown object(s) and write a brief summary of your conclusions.

Materials

- can covered by a sock sealed with tape
- one or more objects that fit in the container
- metric ruler
- balance

SECTION REVIEW

1. List the five main points of Dalton's atomic theory.
2. What chemical laws can be explained by Dalton's theory?

Critical Thinking

3. **ANALYZING INFORMATION** Three compounds containing potassium and oxygen are compared. Analysis shows that for each 1.00 g of O, the compounds have 1.22 g, 2.44 g, and 4.89 g of K, respectively. Show how these data support the law of multiple proportions.

SECTION 2

OBJECTIVES

- Summarize the observed properties of cathode rays that led to the discovery of the electron.
- Summarize the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus.
- List the properties of protons, neutrons, and electrons.
- Define *atom.*

The Structure of the Atom

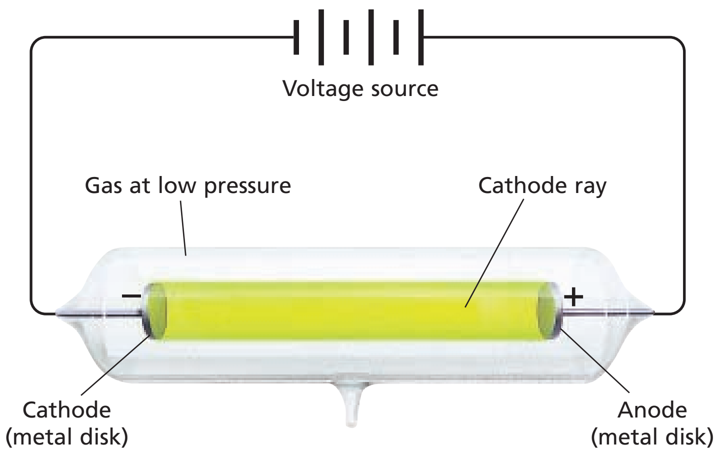
Although John Dalton thought atoms were indivisible, investigators in the late 1800s proved otherwise. As scientific advances allowed a deeper exploration of matter, it became clear that atoms are actually composed of several basic types of smaller particles and that the number and arrangement of these particles within an atom determine that atom's chemical properties. Today we define an **atom** *as the smallest particle of an element that retains the chemical properties of that element.*

All atoms consist of two regions. The *nucleus* is a very small region located at the center of an atom. In every atom, the nucleus is made up of at least one positively charged particle called a *proton* and usually one or more neutral particles called *neutrons.* Surrounding the nucleus is a region occupied by negatively charged particles called *electrons.* This region is very large compared with the size of the nucleus. Protons, neutrons, and electrons are often referred to as *subatomic particles*.

Discovery of the Electron

The first discovery of a subatomic particle resulted from investigations into the relationship between electricity and matter. In the late 1800s, many experiments were performed in which electric current was passed through various gases at low pressures. (Gases at atmospheric pressure don't conduct electricity well.) These experiments were carried out in glass tubes like the one shown in **Figure 4.** Such tubes are known as *cathode-ray tubes*.

FIGURE 4 A simple cathode-ray tube. Particles pass through the tube from the *cathode,* the metal disk connected to the negative terminal of the voltage source, to the *anode,* the metal disk connected to the positive terminal.



Cathode Rays and Electrons

Investigators noticed that when current was passed through a cathode-ray tube, the surface of the tube directly opposite the cathode glowed. They hypothesized that the glow was caused by a stream of particles, which they called a cathode ray. The ray traveled from the cathode to the anode when current was passed through the tube. Experiments devised to test this hypothesis revealed the following observations.

1. Cathode rays were deflected by a magnetic field in the same manner as a wire carrying electric current, which was known to have a negative charge (see **Figure 5**).
2. The rays were deflected away from a negatively charged object.

These observations led to the hypothesis that the particles that compose cathode rays are negatively charged. This hypothesis was strongly supported by a series of experiments carried out in 1897 by the English physicist Joseph John Thomson. In one investigation, he was able to measure the ratio of the charge of cathode-ray particles to their mass. He found that this ratio was always the same, regardless of the metal used to make the cathode or the nature of the gas inside the cathode-ray tube. Thomson concluded that all cathode rays are composed of identical negatively charged particles, which were named electrons.

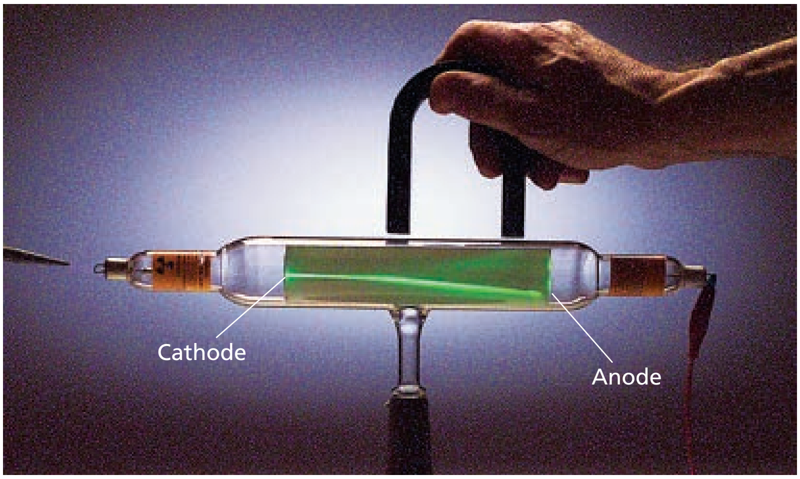


FIGURE 5 A magnet near the cathode-ray tube causes the beam to be deflected. The deflection indicates that the particles in the beam have a negative charge.

Charge and Mass of the Electron

Cathode rays have identical properties regardless of the element used to produce them. Therefore it was concluded that electrons are present in atoms of all elements. Thus, cathode-ray experiments provided evidence that atoms are divisible and that one of the atom's basic constituents is the negatively charged electron. Thomson's experiment also revealed that the electron has a very large charge-to-mass ratio. In 1909, experiments conducted by the American physicist Robert A. Millikan measured the charge of the electron. Scientists used this information and the charge-to-mass ratio of the electron to determine that the mass of the electron is about one two-thousandth the mass of the simplest type of hydrogen atom, which is the smallest atom known. More-accurate experiments conducted since then indicate that the electron has a mass of 9.109×10^{-31} kg, or 1/1837 the mass of the simplest type of hydrogen atom.

Based on what was learned about electrons, two other inferences were made about atomic structure.

1. Because atoms are electrically neutral, they must contain a positive charge to balance the negative electrons.
2. Because electrons have so much less mass than atoms, atoms must contain other particles that account for most of their mass.

Thomson proposed a model for the atom that is called the *plum pudding model* (after the English dessert). He believed that the negative electrons were spread evenly throughout the positive charge of the rest of the atom. This arrangement is similar to that of seeds in a watermelon: the seeds are spread throughout the fruit but do not contribute much to the overall mass. However, shortly thereafter, new experiments disproved this model.

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Topic: Subatomic Particles
Code: HC61473

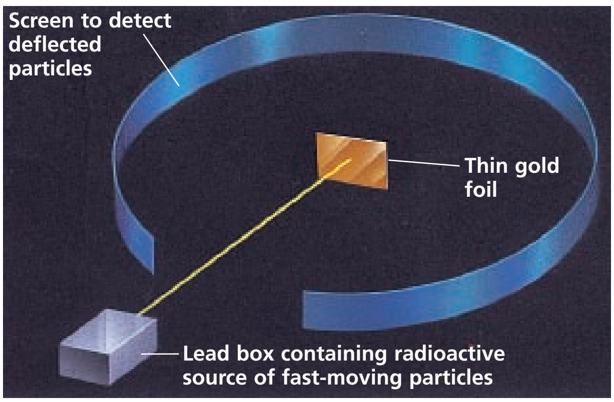
Discovery of the Atomic Nucleus

More detail of the atom's structure was provided in 1911 by New Zealander Ernest Rutherford and his associates Hans Geiger and Ernest Marsden. The scientists bombarded a thin piece of gold foil with fast-moving *alpha particles,* which are positively charged particles with about four times the mass of a hydrogen atom. Geiger and Marsden assumed that mass and charge were uniformly distributed throughout the atoms of the gold foil. They expected the alpha particles to pass through with only a slight deflection, and for the vast majority of the particles, this was the case. However, when the scientists checked for the possibility of wide-angle deflections, they were shocked to find that roughly 1 in 8000 of the alpha particles had actually been deflected back toward the source (see **Figure 6**). As Rutherford later exclaimed, it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue paper and it came back and hit you."

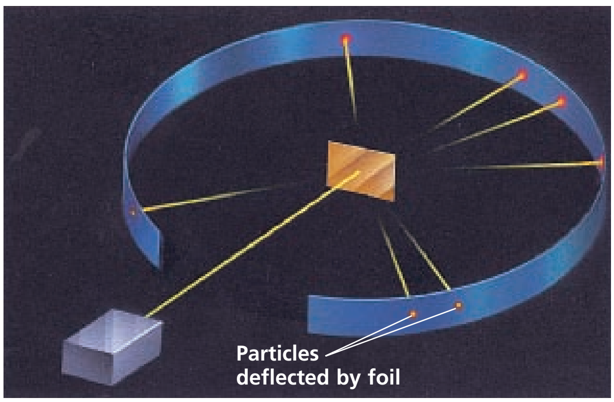
After thinking about the startling result for a few months, Rutherford finally came up with an explanation. He reasoned that the deflected alpha particles must have experienced some powerful force within the atom. And he figured that the source of this force must occupy a very small amount of space because so few of the total number of alpha particles had been affected by it. He concluded that the force must be caused by a very densely packed bundle of matter with a positive electric charge. Rutherford called this positive bundle of matter the nucleus (see **Figure 7**).

Rutherford had discovered that the volume of a nucleus was very small compared with the total volume of an atom. In fact, if the nucleus were the size of a marble, then the size of the atom would be about the size of a football field. But where were the electrons? This question was not answered until Rutherford's student, Niels Bohr, proposed a model in which electrons surrounded the positively charged nucleus as the planets surround the sun. Bohr's model will be discussed in Chapter 4.

FIGURE 6 (a) Geiger and Marsden bombarded a thin piece of gold foil with a narrow beam of alpha particles. (b) Some of the particles were deflected by the gold foil back toward their source.



(a)



(b)

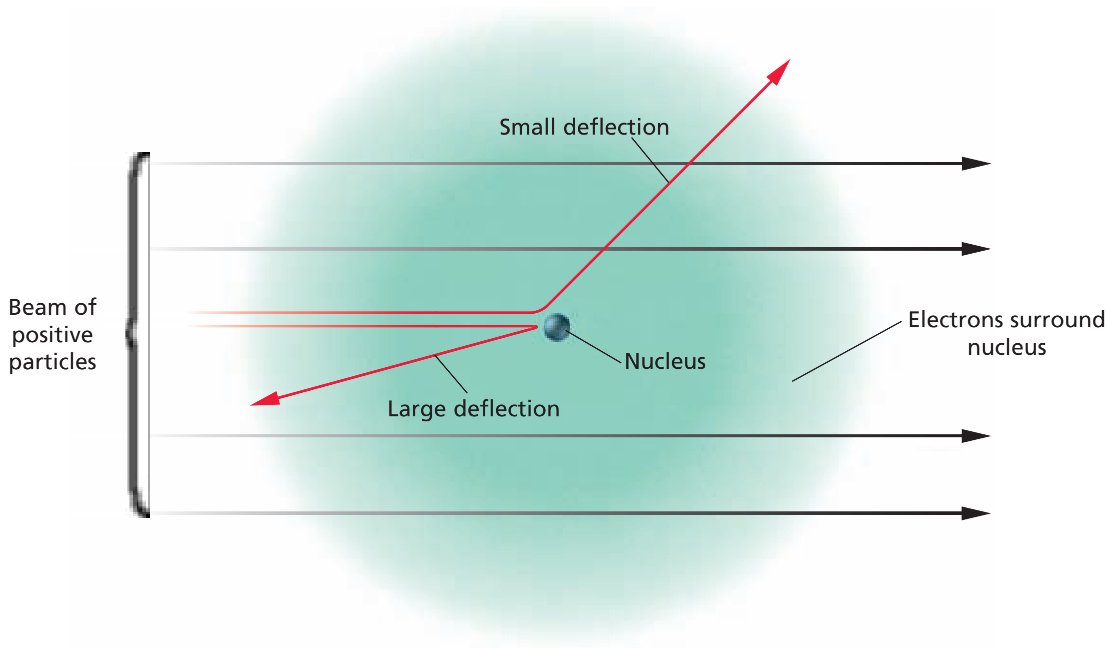


FIGURE 7 Rutherford reasoned that each atom in the gold foil contained a small, dense, positively charged nucleus surrounded by electrons. A small number of the alpha particles directed toward the foil were deflected by the tiny nucleus (red arrows). Most of the particles passed through undisturbed (black arrows).

Composition of the Atomic Nucleus

Except for the nucleus of the simplest type of hydrogen atom (discussed in the next section), all atomic nuclei are made of two kinds of particles, protons and neutrons. A proton has a positive charge equal in magnitude to the negative charge of an electron. Atoms are electrically neutral because they contain equal numbers of protons and electrons. A neutron is electrically neutral.

The simplest hydrogen atom consists of a single-proton nucleus with a single electron moving about it. A proton has a mass of 1.673×10^{-27} kg, which is 1836 times greater than the mass of an electron and 1836/1837, or virtually all, of the mass of the simplest hydrogen atom. All atoms besides the simplest hydrogen atom also have neutrons. The mass of a neutron is 1.675×10^{-27} kg—slightly larger than that of a proton.

The nuclei of atoms of different elements differ in their number of protons and therefore in the amount of positive charge they possess. Thus, the number of protons determines that atom's identity. Physicists have identified other subatomic particles, but particles other than electrons, protons, and neutrons have little effect on the chemical properties of matter. **Table 1** on the next page summarizes the properties of electrons, protons, and neutrons.

Forces in the Nucleus

Generally, particles that have the same electric charge repel one another. Therefore, we would expect a nucleus with more than one proton to be unstable. However, when two protons are extremely close to each other, there is a strong attraction between them. In fact, as many as 83

TABLE 1 *Properties of Subatomic Particles*

Particle	Symbols	Relative electric charge	Mass number	Relative mass (amu*)	Actual mass (kg)
Electron	e^-, ${}^{0}_{-1}e$	-1	0	0.000 5486	9.109×10^{-31}
Proton	p^+, ${}^{1}_{1}\text{H}$	$+1$	1	1.007 276	1.673×10^{-27}
Neutron	n°, ${}^{1}_{0}n$	0	1	1.008 665	1.675×10^{-27}

*1 amu (atomic mass unit) = $1.660\ 540 \times 10^{-27}$ kg

protons can exist close together to help form a stable nucleus. A similar attraction exists when neutrons are very close to each other or when protons and neutrons are very close together. *These short-range proton-neutron, proton-proton, and neutron-neutron forces hold the nuclear particles together and are referred to as* **nuclear forces.**

The Sizes of Atoms

It is convenient to think of the region occupied by the electrons as an electron cloud—a cloud of negative charge. The radius of an atom is the distance from the center of the nucleus to the outer portion of this electron cloud. Because atomic radii are so small, they are expressed using a unit that is more convenient for the sizes of atoms. This unit is the picometer. The abbreviation for the picometer is pm (1 pm = 10^{-12} m = 10^{-10} cm). To get an idea of how small a picometer is, consider that 1 cm is the same fractional part of 10^3 km (about 600 mi) as 100 pm is of 1 cm. Atomic radii range from about 40 to 270 pm. By contrast, the nuclei of atoms have much smaller radii, about 0.001 pm. Nuclei also have incredibly high densities, about 2×10^8 metric tons/cm^3.

SECTION REVIEW

1. Define each of the following:
   a. atom **b.** electron **c.** nucleus **d.** proton **e.** neutron
2. Describe one conclusion made by each of the following scientists that led to the development of the current atomic theory:
   a. Thomson **b.** Millikan **c.** Rutherford
3. Compare the three subatomic particles in terms of location in the atom, mass, and relative charge.
4. Why is the cathode-ray tube in **Figure 4** connected to a vacuum pump?

Critical Thinking

5. **EVALUATING IDEAS** Nuclear forces are said to hold protons and neutrons together. What is it about the composition of the nucleus that requires the concept of nuclear forces?

Counting Atoms

SECTION 3

OBJECTIVES

- Explain what isotopes are.
- Define *atomic number* and *mass number,* and describe how they apply to isotopes.
- Given the identity of a nuclide, determine its number of protons, neutrons, and electrons.
- Define *mole, Avogadro's number,* and *molar mass,* and state how all three are related.
- Solve problems involving mass in grams, amount in moles, and number of atoms of an element.

Consider neon, Ne, the gas used in many illuminated signs. Neon is a minor component of the atmosphere. In fact, dry air contains only about 0.002% neon. And yet there are about 5×10^{17} atoms of neon present in each breath you inhale. In most experiments, atoms are much too small to be measured individually. Chemists can analyze atoms quantitatively, however, by knowing fundamental properties of the atoms of each element. In this section, you will be introduced to some of the basic properties of atoms. You will then discover how to use this information to count the number of atoms of an element in a sample with a known mass. You will also become familiar with the *mole,* a special unit used by chemists to express amounts of particles, such as atoms and molecules.

Atomic Number

All atoms are composed of the same basic particles. Yet all atoms are not the same. Atoms of different elements have different numbers of protons. Atoms of the same element all have the same number of protons. *The* **atomic number** *(Z) of an element is the number of protons of each atom of that element.*

Turn to the inside back cover of this textbook. In the periodic table shown, an element's atomic number is indicated above its symbol. Notice that the elements are placed in order of increasing atomic number. At the top left of the table is hydrogen, H, which has atomic number 1. All atoms of the element hydrogen have one proton. Next in order is helium, He, which has two protons. Lithium, Li, has three protons (see **Figure 8**); beryllium, Be, has four protons; and so on.

The atomic number identifies an element. If you want to know which element has atomic number 47, for example, look at the periodic table. You can see that the element is silver, Ag. All silver atoms have 47 protons. Because atoms are neutral, we know from the atomic number that all silver atoms must also have 47 electrons.

3
Li
Lithium
6.941
$[He]2s^1$

FIGURE 8 The atomic number in this periodic table entry reveals that an atom of lithium has three protons in its nucleus.

Isotopes

The simplest atoms are those of hydrogen. All hydrogen atoms have only one proton. However, like many naturally occurring elements, hydrogen atoms can have different numbers of neutrons.

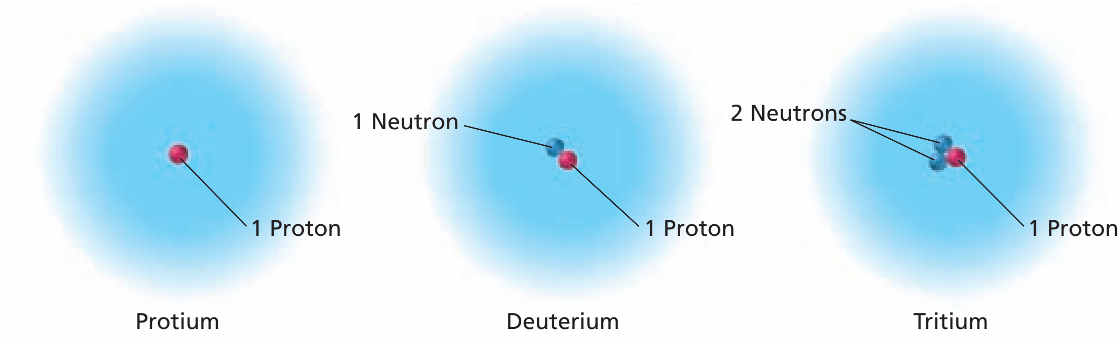


FIGURE 9 The nuclei of different isotopes of the same element have the same number of protons but different numbers of neutrons. This is illustrated above by the three isotopes of hydrogen.

Three types of hydrogen atoms are known. The most common type of hydrogen is sometimes called *protium.* It accounts for 99.9885% of the hydrogen atoms found on Earth. The nucleus of a protium atom consists of one proton only, and it has one electron moving about it. There are two other known forms of hydrogen. One is called *deuterium,* which accounts for 0.0115% of Earth's hydrogen atoms. Each deuterium atom has a nucleus with one proton and one neutron. The third form of hydrogen is known as *tritium,* which is radioactive. It exists in very small amounts in nature, but it can be prepared artificially. Each tritium atom has one proton, two neutrons, and one electron.

Protium, deuterium, and tritium are isotopes of hydrogen. **Isotopes** *are atoms of the same element that have different masses.* The isotopes of a particular element all have the same number of protons and electrons but different numbers of neutrons. In all three isotopes of hydrogen, the positive charge of the single proton is balanced by the negative charge of the electron. Most of the elements consist of mixtures of isotopes. Tin has 10 stable isotopes, for example, the most of any element.

Mass Number

Identifying an isotope requires knowing both the name or atomic number of the element and the mass of the isotope. *The* **mass number** *is the total number of protons and neutrons that make up the nucleus of an isotope.* The three isotopes of hydrogen described earlier have mass numbers 1, 2, and 3, as shown in **Table 2.**

TABLE 2 *Mass Numbers of Hydrogen Isotopes*

	Atomic number (number of protons)	Number of neutrons	Mass number (protons + neutrons)
Protium	1	0	$1 + 0 = 1$
Deuterium	1	1	$1 + 1 = 2$
Tritium	1	2	$1 + 2 = 3$

Designating Isotopes

The isotopes of hydrogen are unusual in that they have distinct names. Isotopes are usually identified by specifying their mass number. There are two methods for specifying isotopes. In the first method, the mass number is written with a hyphen after the name of the element. Tritium, for example, is written as hydrogen-3. We will refer to this method as *hyphen notation.* The uranium isotope used as fuel for nuclear power plants has a mass number of 235 and is therefore known as uranium-235. The second method shows the composition of a nucleus using the isotope's *nuclear symbol*. For example, uranium-235 is written as $^{235}_{92}U$. The superscript indicates the mass number (protons + neutrons) and the subscript indicates the atomic number (number of protons). The number of neutrons is found by subtracting the atomic number from the mass number.

mass number − atomic number = number of neutrons
235 (protons + neutrons) − 92 protons = 143 neutrons

Thus, a uranium-235 nucleus is made up of 92 protons and 143 neutrons.

Table 3 gives the names, symbols, and compositions of the isotopes of hydrogen and helium. **Nuclide** *is a general term for a specific isotope of an element.* We could say that **Table 3** lists the compositions of five different nuclides, three hydrogen nuclides and two helium nuclides.

TABLE 3 Isotopes of Hydrogen and Helium

Isotope	Nuclear symbol	Number of protons	Number of electrons	Number of neutrons
Hydrogen-1 (protium)	$^{1}_{1}H$	1	1	0
Hydrogen-2 (deuterium)	$^{2}_{1}H$	1	1	1
Hydrogen-3 (tritium)	$^{3}_{1}H$	1	1	2
Helium-3	$^{3}_{2}He$	2	2	1
Helium-4	$^{4}_{2}He$	2	2	2

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Topic: Isotopes
Code: HC60820

SAMPLE PROBLEM A

How many protons, electrons, and neutrons are there in an atom of chlorine-37?

SOLUTION

1 *ANALYZE* **Given:** name and mass number of chlorine-37
Unknown: numbers of protons, electrons, and neutrons

2 *PLAN* atomic number = number of protons = number of electrons
mass number = number of neutrons + number of protons

3 ***COMPUTE*** The mass number of chlorine-37 is 37. Consulting the periodic table reveals that chlorine's atomic number is 17. The number of neutrons can be found by subtracting the atomic number from the mass number.

mass number of chlorine-37 – atomic number of chlorine = number of neutrons in chlorine-37

mass number – atomic number = 37 (protons plus neutrons) – 17 protons = 20 neutrons

An atom of chlorine-37 is made up of 17 electrons, 17 protons, and 20 neutrons.

4 ***EVALUATE*** The number of protons in a neutral atom equals the number of electrons. And the sum of the protons and neutrons equals the given mass number.

PRACTICE *Answers in Appendix E*

1. How many protons, electrons, and neutrons make up an atom of bromine-80?
2. Write the nuclear symbol for carbon-13.
3. Write the hyphen notation for the isotope with 15 electrons and 15 neutrons.

extension

Go to **go.hrw.com** for more practice problems that ask you to work with numbers of subatomic particles.

Keyword: HC6ATMX

Relative Atomic Masses

Masses of atoms expressed in grams are very small. As we shall see, an atom of oxygen-16, for example, has a mass of 2.656×10^{-23} g. For most chemical calculations it is more convenient to use *relative* atomic masses. As you read in Chapter 2, scientists use standards of measurement that are constant and are the same everywhere. In order to set up a relative scale of atomic mass, one atom has been arbitrarily chosen as the standard and assigned a mass value. The masses of all other atoms are expressed in relation to this defined standard.

The standard used by scientists to compare units of atomic mass is the carbon-12 atom. It has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12 amu. *One* **atomic mass unit,** *or 1 amu, is exactly 1/12 the mass of a carbon-12 atom.* The atomic mass of any other atom is determined by comparing it with the mass of the carbon-12 atom. The hydrogen-1 atom has an atomic mass of *about* 1/12 that of the carbon-12 atom, or about 1 amu. The precise value of the atomic mass of a hydrogen-1 atom is 1.007 825 amu. An oxygen-16 atom has about 16/12 (or 4/3) the mass of a carbon-12 atom. Careful measurements show the atomic mass of oxygen-16 to be 15.994 915 amu. The mass of a magnesium-24 atom is found to be slightly less than twice that of a carbon-12 atom. Its atomic mass is 23.985 042 amu.

Some additional examples of the atomic masses of the naturally occurring isotopes of several elements are given in **Table 4** on the next page. Isotopes of an element may occur naturally, or they may be made in the laboratory (*artificial isotopes*). *Although isotopes have different masses, they do not differ significantly in their chemical behavior*.

The masses of subatomic particles can also be expressed on the atomic mass scale (see **Table 1**). The mass of the electron is 0.000 5486 amu, that of the proton is 1.007 276 amu, and that of the neutron is 1.008 665 amu. Note that the proton and neutron masses are close to but not equal to 1 amu. You have learned that the mass number is the total number of protons and neutrons that make up the nucleus of an atom. You can now see that the mass number and relative atomic mass of a given nuclide are quite close to each other. They are not identical because the proton and neutron masses deviate slightly from 1 amu and the atomic masses include electrons. Also, as you will read in Chapter 21, a small amount of mass is changed to energy in the creation of a nucleus from its protons and neutrons.

HISTORICAL CHEMISTRY

Discovery of Element 43

The discovery of element 43, technetium, is credited to Carlo Perrier and Emilio Segrè, who artificially produced it in 1937. However, in 1925, a German chemist named Ida Tacke reported the discovery of element 43, which she called masurium, in niobium ores. At the time, her discovery was not accepted because it was thought technetium could not occur naturally. Recent studies confirm that Tacke and coworkers probably did discover element 43.

Average Atomic Masses of Elements

Most elements occur naturally as mixtures of isotopes, as indicated in **Table 4.** The percentage of each isotope in the naturally occurring element on Earth is nearly always the same, no matter where the element is found. The percentage at which each of an element's isotopes occurs in nature is taken into account when calculating the element's average atomic mass. **Average atomic mass** *is the weighted average of the atomic masses of the naturally occurring isotopes of an element.*

The following is a simple example of how to calculate a *weighted average.* Suppose you have a box containing two sizes of marbles. If 25% of the marbles have masses of 2.00 g each and 75% have masses of 3.00 g each, how is the weighted average calculated? You could count the number of each type of marble, calculate the total mass of the mixture, and divide by the total number of marbles. If you had 100 marbles, the calculations would be as follows.

$$25 \text{ marbles} \times 2.00 \text{ g} = 50 \text{ g}$$
$$75 \text{ marbles} \times 3.00 \text{ g} = 225 \text{ g}$$

Adding these masses gives the total mass of the marbles.

$$50 \text{ g} + 225 \text{ g} = 275 \text{ g}$$

Dividing the total mass by 100 gives an average marble mass of 2.75 g.

A simpler method is to multiply the mass of each marble by the decimal fraction representing its percentage in the mixture. Then add the products.

$$25\% = 0.25 \qquad 75\% = 0.75$$
$$(2.00 \text{ g} \times 0.25) + (3.00 \text{ g} \times 0.75) = 2.75 \text{ g}$$

TABLE 4 Atomic Masses and Abundances of Several Naturally Occurring Isotopes

Isotope	Mass number	Percentage natural abundance	Atomic mass (amu)	Average atomic mass of element (amu)
Hydrogen-1	1	99.9885	1.007 825	1.007 94
Hydrogen-2	2	0.0115	2.014 102	
Carbon-12	12	98.93	12 (by definition)	
Carbon-13	13	1.07	13.003 355	12.0107
Oxygen-16	16	99.757	15.994 915	
Oxygen-17	17	0.038	16.999 132	15.9994
Oxygen-18	18	0.205	17.999 160	
Copper-63	63	69.15	62.929 601	63.546
Copper-65	65	30.85	64.927 794	
Cesium-133	133	100	132.905 447	132.905
Uranium-234	234	0.0054	234.040 945	
Uranium-235	235	0.7204	235.043 922	238.029
Uranium-238	238	99.2742	238.050 784	

Calculating Average Atomic Mass

The average atomic mass of an element depends on both the mass and the relative abundance of each of the element's isotopes. For example, naturally occurring copper consists of 69.15% copper-63, which has an atomic mass of 62.929 601 amu, and 30.85% copper-65, which has an atomic mass of 64.927 794 amu. The average atomic mass of copper can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.

$$0.6915 \times 62.929\,601 \text{ amu} + 0.3085 \times 64.927\,794 \text{ amu} = 63.55 \text{ amu}$$

The calculated average atomic mass of naturally occurring copper is 63.55 amu.

The average atomic mass is included for the elements listed in **Table 4.** As illustrated in the table, most atomic masses are known to four or more significant figures. *In this book, an element's atomic mass is usually rounded to two decimal places before it is used in a calculation.*

Relating Mass to Numbers of Atoms

The relative atomic mass scale makes it possible to know how many atoms of an element are present in a sample of the element with a measurable mass. Three very important concepts—the mole, Avogadro's number, and molar mass—provide the basis for relating masses in grams to numbers of atoms.

The Mole

The mole is the SI unit for amount of substance. *A* **mole** *(abbreviated* mol*) is the amount of a substance that contains as many particles as there are atoms in exactly 12 g of carbon-12.* The mole is a counting unit, just like a dozen is. We don't usually order 12 or 24 ears of corn; we order one dozen or two dozen. Similarly, a chemist may want 1 mol of carbon, or 2 mol of iron, or 2.567 mol of calcium. In the sections that follow, you will see how the mole relates to masses of atoms and compounds.

Avogadro's Number

The number of particles in a mole has been experimentally determined in a number of ways. The best modern value is $6.022\ 141\ 79 \times 10^{23}$. This means that exactly 12 g of carbon-12 contains $6.022\ 141\ 79 \times 10^{23}$ carbon-12 atoms. The number of particles in a mole is known as Avogadro's number, named for the nineteenth-century Italian scientist Amedeo Avogadro, whose ideas were crucial in explaining the relationship between mass and numbers of atoms. **Avogadro's number**—*$6.022\ 141\ 79 \times 10^{23}$—is the number of particles in exactly one mole of a pure substance.* For most purposes, Avogadro's number is rounded to 6.022×10^{23}.

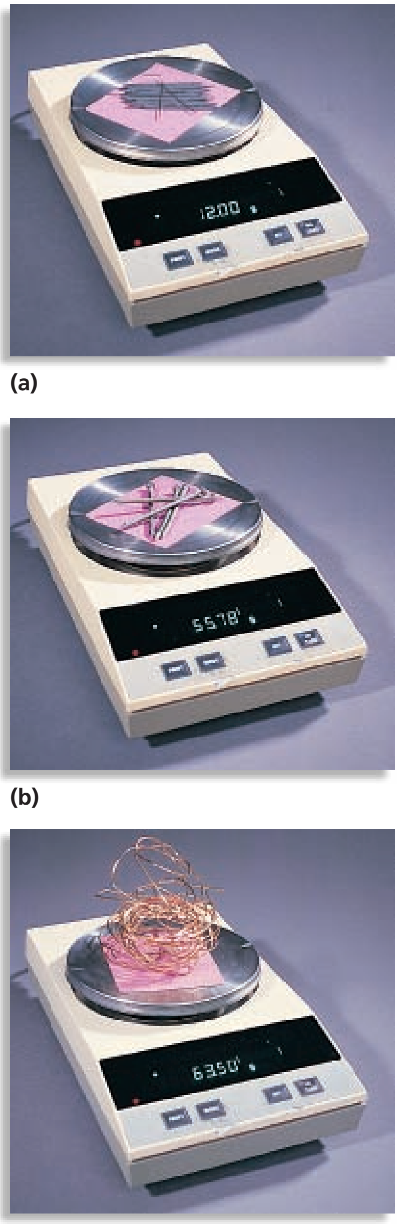
To get a sense of how large Avogadro's number is, consider the following: If every person living on Earth (6 billion people) worked to count the atoms in one mole of an element, and if each person counted continuously at a rate of one atom per second, it would take about 3 million years for all the atoms to be counted.

Molar Mass

An alternative definition of *mole* is the amount of a substance that contains Avogadro's number of particles. Can you figure out the approximate mass of one mole of helium atoms? You know that a mole of carbon-12 atoms has a mass of exactly 12 g and that a carbon-12 atom has an atomic mass of 12 amu. The atomic mass of a helium atom is 4.00 amu, which is about one-third the mass of a carbon-12 atom. It follows that a mole of helium atoms will have about one-third the mass of a mole of carbon-12 atoms. Thus, one mole of helium has a mass of about 4.00 g.

The mass of one mole of a pure substance is called the **molar mass** *of that substance.* Molar mass is usually written in units of g/mol. The molar mass of an element is numerically equal to the atomic mass of the element in atomic mass units (which can be found in the periodic table). For example, the molar mass of lithium, Li, is 6.94 g/mol, while the molar mass of mercury, Hg, is 200.59 g/mol (rounding each value to two decimal places).

The molar mass of an element contains one mole of atoms. For example, 4.00 g of helium, 6.94 g of lithium, and 200.59 g of mercury all contain a mole of atoms. **Figure 10** shows molar masses of three common elements.



(a)

(b)

(c)

FIGURE 10 Shown is approximately one molar mass of each of three elements: (a) carbon (graphite), (b) iron (nails), and (c) copper (wire).

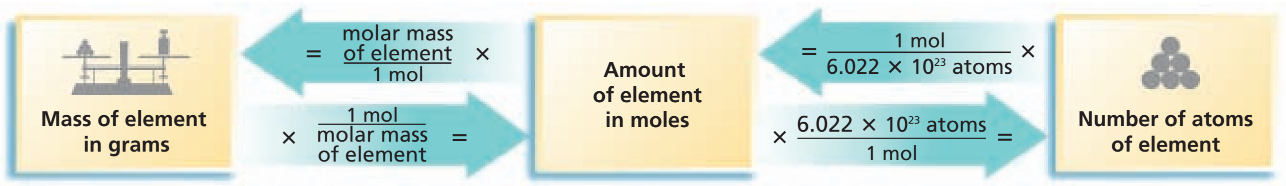


FIGURE 11 The diagram shows the relationship between mass in grams, amount in moles, and number of atoms of an element in a sample.

Gram/Mole Conversions

Chemists use molar mass as a conversion factor in chemical calculations. For example, the molar mass of helium is 4.00 g He/mol He. To find how many grams of helium there are in two moles of helium, multiply by the molar mass.

$$2.00\ \cancel{\text{mol He}} \times \frac{4.00\ \text{g He}}{1\ \cancel{\text{mol He}}} = 8.00\ \text{g He}$$

Figure 11 shows how to use molar mass, moles, and Avogadro's number to relate mass in grams, amount in moles, and number of atoms of an element.

SAMPLE PROBLEM B *For more help, go to the Math Tutor at the end of this chapter.*

What is the mass in grams of 3.50 mol of the element copper, Cu?

SOLUTION

1 ***ANALYZE***

Given: 3.50 mol Cu
Unknown: mass of Cu in grams

2 ***PLAN***

$$\text{amount of Cu in moles} \longrightarrow \text{mass of Cu in grams}$$

According to **Figure 11,** the mass of an element in grams can be calculated by multiplying the amount of the element in moles by the element's molar mass.

$$\text{moles Cu} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

3 ***COMPUTE***

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$3.50\ \cancel{\text{mol Cu}} \times \frac{63.55\ \text{g Cu}}{1\ \cancel{\text{mol Cu}}} = 222\ \text{g Cu}$$

4 ***EVALUATE***

Because the amount of copper in moles was given to three significant figures, the answer was rounded to three significant figures. The size of the answer is reasonable because it is somewhat more than 3.5 times 60.

PRACTICE

Answers in Appendix E

1. What is the mass in grams of 2.25 mol of the element iron, Fe?
2. What is the mass in grams of 0.375 mol of the element potassium, K?
3. What is the mass in grams of 0.0135 mol of the element sodium, Na?
4. What is the mass in grams of 16.3 mol of the element nickel, Ni?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert from amount in moles to mass.

Keyword: HC6ATMX

SAMPLE PROBLEM C

For more help, go to the Math Tutor at the end of this chapter.

A chemist produced 11.9 g of aluminum, Al. How many moles of aluminum were produced?

SOLUTION

1 ANALYZE

Given: 11.9 g Al

Unknown: amount of Al in moles

2 PLAN

$$\text{mass of Al in grams} \longrightarrow \text{amount of Al in moles}$$

As shown in **Figure 11,** amount in moles can be obtained by *dividing* mass in grams by molar mass, which is mathematically the same as *multiplying* mass in grams by the *reciprocal* of molar mass.

$$\text{grams Al} \times \frac{\text{moles Al}}{\text{grams Al}} = \text{moles Al}$$

3 COMPUTE

The molar mass of aluminum from the periodic table is rounded to 26.98 g/mol.

$$11.9\ \cancel{\text{g Al}} \times \frac{1\ \text{mol Al}}{26.98\ \cancel{\text{g Al}}} = 0.441\ \text{mol Al}$$

4 EVALUATE

The answer is correctly given to three significant figures. The answer is reasonable because 11.9 g is somewhat less than half of 26.98 g.

PRACTICE

Answers in Appendix E

1. How many moles of calcium, Ca, are in 5.00 g of calcium?
2. How many moles of gold, Au, are in 3.60×10^{-5} g of gold?
3. How many moles of zinc, Zn, are in 0.535 g of zinc?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert from mass to amount in moles.

Keyword: HC6ATMX

Conversions with Avogadro's Number

Figure 11 shows that Avogadro's number can be used to find the number of atoms of an element from the amount in moles or to find the amount of an element in moles from the number of atoms. While these types of problems are less common in chemistry than converting between amount in moles and mass in grams, they are useful in demonstrating the meaning of Avogadro's number. Note that in these calculations, Avogadro's number is expressed in units of atoms per mole.

SAMPLE PROBLEM D *For more help, go to the Math Tutor at the end of this chapter.*

How many moles of silver, Ag, are in 3.01×10^{23} atoms of silver?

SOLUTION

1 *ANALYZE* **Given:** 3.01×10^{23} atoms of Ag
Unknown: amount of Ag in moles

2 *PLAN*

$$\text{number of atoms of Ag} \longrightarrow \text{amount of Ag in moles}$$

From **Figure 11,** we know that number of atoms is converted to amount in moles by dividing by Avogadro's number. This is equivalent to multiplying numbers of atoms by the reciprocal of Avogadro's number.

$$\text{Ag atoms} \times \frac{\text{moles Ag}}{\text{Avogadro's number of Ag atoms}} = \text{moles Ag}$$

3 *COMPUTE*

$$3.01 \times 10^{23}\ \cancel{\text{Ag atoms}} \times \frac{1\ \text{mol Ag}}{6.022 \times 10^{23}\ \cancel{\text{Ag atoms}}} = 0.500\ \text{mol Ag}$$

4 *EVALUATE* The answer is correct—units cancel correctly and the number of atoms is one-half of Avogadro's number.

PRACTICE *Answers in Appendix E*

1. How many moles of lead, Pb, are in 1.50×10^{12} atoms of lead?
2. How many moles of tin, Sn, are in 2500 atoms of tin?
3. How many atoms of aluminum, Al, are in 2.75 mol of aluminum?

extension
Go to **go.hrw.com** for more practice problems that ask you to convert between atoms and moles.
Keyword: HC6ATMX

SAMPLE PROBLEM E *For more help, go to the Math Tutor at the end of this chapter.*

What is the mass in grams of 1.20×10^{8} atoms of copper, Cu?

SOLUTION

1 *ANALYZE* **Given:** 1.20×10^{8} atoms of Cu
Unknown: mass of Cu in grams

2 PLAN

number of atoms of Cu $\longrightarrow$ amount of Cu in moles $\longrightarrow$ mass of Cu in grams

As indicated in **Figure 11,** the given number of atoms must first be converted to amount in moles by dividing by Avogadro's number. Amount in moles is then multiplied by molar mass to yield mass in grams.

$$\text{Cu atoms} \times \frac{\text{moles Cu}}{\text{Avogadro's number of Cu atoms}} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

3 COMPUTE

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$1.20 \times 10^8 \cancel{\text{Cu atoms}} \times \frac{1 \cancel{\text{mol Cu}}}{6.022 \times 10^{23} \cancel{\text{Cu atoms}}} \times \frac{63.55 \text{ g Cu}}{1 \cancel{\text{mol Cu}}} = 1.27 \times 10^{-14} \text{ g Cu}$$

4 EVALUATE

Units cancel correctly to give the answer in grams. The size of the answer is reasonable—10^8 has been divided by about 10^{24} and multiplied by about 10^2.

PRACTICE

Answers in Appendix E

1. What is the mass in grams of 7.5×10^{15} atoms of nickel, Ni?
2. How many atoms of sulfur, S, are in 4.00 g of sulfur?
3. What mass of gold, Au, contains the same number of atoms as 9.0 g of aluminum, Al?

extension

Go to **go.hrw.com** for more practice problems that ask you to convert among atoms, grams, and moles.

Keyword: HC6ATMX

SECTION REVIEW

1. Define each of the following:
   - **a.** atomic number
   - **b.** mass number
   - **c.** relative atomic mass
   - **d.** average atomic mass
   - **e.** mole
   - **f.** Avogadro's number
   - **g.** molar mass
   - **h.** isotope
2. Determine the number of protons, electrons, and neutrons in each of the following isotopes:
   - **a.** sodium-23
   - **b.** calcium-40
   - **c.** $^{64}_{29}Cu$
   - **d.** $^{108}_{47}Ag$
3. Write the nuclear symbol and hyphen notation for each of the following isotopes:
   - **a.** mass number of 28 and atomic number of 14
   - **b.** 26 protons and 30 neutrons
4. To two decimal places, what is the relative atomic mass and the molar mass of the element potassium, K?
5. Determine the mass in grams of the following:
   - **a.** 2.00 mol N
   - **b.** 3.01×10^{23} atoms Cl
6. Determine the amount in moles of the following:
   - **a.** 12.15 g Mg
   - **b.** 1.50×10^{23} atoms F

Critical Thinking

7. **ANALYZING DATA** Beaker A contains 2.06 mol of copper, and Beaker B contains 222 grams of silver. Which beaker contains the larger mass? Which beaker has the larger number of atoms?

CHAPTER HIGHLIGHTS

The Atom: From Philosophical Idea to Scientific Theory

Vocabulary

law of conservation of mass
law of definite proportions
law of multiple proportions

- The idea of atoms has been around since the time of the ancient Greeks. In the nineteenth century, John Dalton proposed a scientific theory of atoms that can still be used to explain properties of most chemicals today.
- Matter and its mass cannot be created or destroyed in chemical reactions.
- The mass ratios of the elements that make up a given compound are always the same, regardless of how much of the compound there is or how it was formed.
- If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element can be expressed as a ratio of small whole numbers.

The Structure of the Atom

Vocabulary

atom
nuclear forces

- Cathode-ray tubes supplied evidence of the existence of electrons, which are negatively charged subatomic particles that have relatively little mass.
- Rutherford found evidence for the existence of the atomic nucleus by bombarding gold foil with a beam of positively charged particles.
- Atomic nuclei are composed of protons, which have an electric charge of +1, and (in all but one case) neutrons, which have no electric charge.
- Atomic nuclei have radii of about 0.001 pm (pm = picometers; 1 pm $\times$ 10^{-12} m), and atoms have radii of about 40–270 pm.

Counting Atoms

Vocabulary

atomic number
isotope
mass number
nuclide
atomic mass unit
average atomic mass
mole
Avogadro's number
molar mass

- The atomic number of an element is equal to the number of protons of an atom of that element.
- The mass number is equal to the total number of protons and neutrons that make up the nucleus of an atom of that element.
- The relative atomic mass unit (amu) is based on the carbon-12 atom and is a convenient unit for measuring the mass of atoms. It equals $1.660\,540 \times 10^{-24}$ g.
- The average atomic mass of an element is found by calculating the weighted average of the atomic masses of the naturally occurring isotopes of the element.
- Avogadro's number is equal to approximately 6.022×10^{23}. A sample that contains a number of particles equal to Avogadro's number contains a mole of those particles.

CHAPTER REVIEW

For more practice, go to the Problem Bank in Appendix D.

The Atom: From Philosophical Idea to Scientific Theory

SECTION 1 REVIEW

1. Explain each of the following in terms of Dalton's atomic theory:
a. the law of conservation of mass
b. the law of definite proportions
c. the law of multiple proportions

2. According to the law of conservation of mass, if element A has an atomic mass of 2 mass units and element B has an atomic mass of 3 mass units, what mass would be expected for compound AB? for compound A_2B_3?

The Structure of the Atom

SECTION 2 REVIEW

3. a. What is an atom?
b. What two regions make up all atoms?

4. Describe at least four properties of electrons that were determined based on the experiments of Thomson and Millikan.

5. Summarize Rutherford's model of the atom, and explain how he developed this model based on the results of his famous gold-foil experiment.

6. What number uniquely identifies an element?

Counting Atoms

SECTION 3 REVIEW

7. a. What are isotopes?
b. How are the isotopes of a particular element alike?
c. How are they different?

8. Copy and complete the following table concerning the three isotopes of silicon, Si. (Hint: See Sample Problem A.)

Isotope	Number of protons	Number of electrons	Number of neutrons
Si-28			
Si-29			
Si-30			

9. a. What is the atomic number of an element?
b. What is the mass number of an isotope?
c. In the nuclear symbol for deuterium, ${}^{2}_{1}H$, identify the atomic number and the mass number.

10. What is a nuclide?

11. Use the periodic table and the information that follows to write the hyphen notation for each isotope described.
a. atomic number = 2, mass number = 4
b. atomic number = 8, mass number = 16
c. atomic number = 19, mass number = 39

12. a. What nuclide is used as the standard in the relative scale for atomic masses?
b. What is its assigned atomic mass?

13. What is the atomic mass of an atom if its mass is approximately equal to the following?
a. $\frac{1}{3}$ that of carbon-12
b. 4.5 times as much as carbon-12

14. a. What is the definition of a *mole*?
b. What is the abbreviation for *mole*?
c. How many particles are in one mole?
d. What name is given to the number of particles in a mole?

15. a. What is the molar mass of an element?
b. To two decimal places, write the molar masses of carbon, neon, iron, and uranium.

16. Suppose you have a sample of an element.
a. How is the mass in grams of the element converted to amount in moles?
b. How is the mass in grams of the element converted to number of atoms?

PRACTICE PROBLEMS

17. What is the mass in grams of each of the following? (Hint: See Sample Problems B and E.)
a. 1.00 mol Li
b. 1.00 mol Al
c. 1.00 molar mass Ca
d. 1.00 molar mass Fe
e. 6.022×10^{23} atoms C
f. 6.022×10^{23} atoms Ag

18. How many moles of atoms are there in each of the following? (Hint: See Sample Problems C and D.)
a. 6.022×10^{23} atoms Ne
b. 3.011×10^{23} atoms Mg
c. 3.25×10^{5} g Pb
d. 4.50×10^{-12} g O

19. Three isotopes of argon occur in nature—$^{36}_{18}Ar$, $^{38}_{18}Ar$, and $^{40}_{18}Ar$. Calculate the average atomic mass of argon to two decimal places, given the following relative atomic masses and abundances of each of the isotopes: argon-36 (35.97 amu; 0.337%), argon-38 (37.96 amu; 0.063%), and argon-40 (39.96 amu; 99.600%).

20. Naturally occurring boron is 80.20% boron-11 (atomic mass = 11.01 amu) and 19.80% of some other isotopic form of boron. What must the atomic mass of this second isotope be in order to account for the 10.81 amu average atomic mass of boron? (Write the answer to two decimal places.)

21. How many atoms are there in each of the following?
a. 1.50 mol Na
b. 6.755 mol Pb
c. 7.02 g Si

22. What is the mass in grams of each of the following?
a. 3.011×10^{23} atoms F
b. 1.50×10^{23} atoms Mg
c. 4.50×10^{12} atoms Cl
d. 8.42×10^{18} atoms Br
e. 25 atoms W
f. 1 atom Au

23. Determine the number of atoms in each of the following:
a. 5.40 g B
b. 0.250 mol S
c. 0.0384 mol K
d. 0.025 50 g Pt
e. 1.00×10^{-10} g Au

MIXED REVIEW

24. Determine the mass in grams of each of the following:
a. 3.00 mol Al
b. 2.56×10^{24} atoms Li
c. 1.38 mol N
d. 4.86×10^{24} atoms Au
e. 6.50 mol Cu
f. 2.57×10^{8} mol S
g. 1.05×10^{18} atoms Hg

25. Copy and complete the following table concerning the properties of subatomic particles.

Particle	Symbol	Mass number	Actual mass	Relative charge
Electron				
Proton				
Neutron				

26. a. How is an atomic mass unit (amu) related to the mass of one carbon-12 atom?
b. What is the relative atomic mass of an atom?

27. a. What is the nucleus of an atom?
b. Who is credited with the discovery of the atomic nucleus?
c. Identify the two kinds of particles that make up the nucleus.

28. How many moles of atoms are there in each of the following?
a. 40.1 g Ca
b. 11.5 g Na
c. 5.87 g Ni
d. 150 g S
e. 2.65 g Fe
f. 0.007 50 g Ag
g. 2.25×10^{25} atoms Zn
h. 50 atoms Ba

29. State the law of multiple proportions, and give an example of two compounds that illustrate the law.

30. What is the approximate atomic mass of an atom if its mass is
a. 12 times that of carbon-12?
b. $\frac{1}{2}$ that of carbon-12?

31. What is an electron?

CRITICAL THINKING

32. **Organizing Ideas** Using two chemical compounds as an example, describe the difference between the law of definite proportions and the law of multiple proportions.

33. **Constructing Models** As described in Section 2, the structure of the atom was determined from observations made in painstaking experimental research. Suppose a series of experiments revealed that when an electric current is passed through gas at low pressure, the surface of the cathode-ray tube opposite the anode glows. In addition, a paddle wheel placed in the tube rolls from the anode toward the cathode when the current is on.

a. In which direction do particles pass through the gas?

b. What charge do the particles possess?

34. **Analyzing Data** Osmium is the element with the greatest density, 22.58 g/cm^3. How does the density of osmium compare to the density of a typical nucleus of 2×10^8 metric tons/cm^3? (1 metric ton = 1000 kg)

USING THE HANDBOOK

35. Group 14 of the *Elements Handbook* describes the reactions that produce CO and CO_2. Review this section to answer the following:

a. When a fuel burns, what determines whether CO or CO_2 will be produced?

b. What happens in the body if hemoglobin picks up CO?

c. Why is CO poisoning most likely to occur in homes that are well sealed during cold winter months?

RESEARCH & WRITING

36. Prepare a report on the series of experiments conducted by Sir James Chadwick that led to the discovery of the neutron.

37. Write a report on the contributions of Amedeo Avogadro that led to the determination of the value of Avogadro's number.

38. Trace the development of the electron microscope, and cite some of its many uses.

39. The study of atomic structure and the nucleus produced a new field of medicine called *nuclear medicine.* Describe the use of radioactive tracers to detect and treat diseases.

ALTERNATIVE ASSESSMENT

40. Observe a cathode-ray tube in operation, and write a description of your observations.

41. **Performance Assessment** Using colored clay, build a model of the nucleus of each of carbon's three naturally occurring isotopes: carbon-12, carbon-13, and carbon-14. Specify the number of electrons that would surround each nucleus.

extension

Graphing Calculator **Calculating Numbers of Protons, Electrons, and Neutrons**

Go to **go.hrw.com** for a graphing calculator exercise that asks you to calculate numbers of protons, electrons, and neutrons.

Keyword: HC6ATMX

Math Tutor CONVERSION FACTORS

Most calculations in chemistry require that all measurements of the same quantity (mass, length, volume, temperature, and so on) be expressed in the same unit. To change the units of a quantity, you can multiply the quantity by a conversion factor. With SI units, such conversions are easy because units of the same quantity are related by multiples of 10, 100, 1000, or 1 million. Suppose you want to convert a given amount in milliliters to liters. You can use the relationship 1 L = 1000 mL. From this relationship, you can derive the following conversion factors.

$$\frac{1000 \text{ mL}}{1 \text{ L}} \text{ and } \frac{1 \text{ L}}{1000 \text{ mL}}$$

The correct strategy is to multiply the given amount (in mL) by the conversion factor that allows milliliter units to cancel out and liter units to remain. Using the second conversion factor will give you the units you want.

These conversion factors are based on an exact definition (1000 mL = 1 L exactly), so significant figures do not apply to these factors. The number of significant figures in a converted measurement depends on the certainty of the measurement you start with.

SAMPLE 1

A sample of aluminum has a mass of 0.087 g. What is the sample's mass in milligrams?

Based on SI prefixes, you know that 1 g = 1000 mg. Therefore, the possible conversion factors are

$$\frac{1000 \text{ mg}}{1 \text{ g}} \text{ and } \frac{1 \text{ g}}{1000 \text{ mg}}$$

The first conversion factor cancels grams, leaving milligrams.

$$0.087 \cancel{\text{g}} \times \frac{1000 \text{ mg}}{1 \cancel{\text{g}}} = 87 \text{ mg}$$

Notice that the values 0.087 g and 87 mg each have two significant figures.

SAMPLE 2

A sample of a mineral has 4.08×10^{-5} mol of vanadium per kilogram of mass. How many micromoles of vanadium per kilogram does the mineral contain?

The prefix *micro-* specifies $\frac{1}{1\,000\,000}$ or 1×10^{-6} of the base unit.

So, 1 μmol = 1×10^{-6} mol. The possible conversion factors are

$$\frac{1 \ \mu\text{mol}}{1 \times 10^{-6} \text{ mol}} \text{ and } \frac{1 \times 10^{-6} \text{ mol}}{1 \ \mu\text{mol}}$$

The first conversion factor will allow moles to cancel and micromoles to remain.

$$4.08 \times 10^{-5} \cancel{\text{mol}} \times \frac{1 \ \mu\text{mol}}{1 \times 10^{-6} \cancel{\text{mol}}} = 40.8 \ \mu\text{mol}$$

Notice that the values 4.08×10^{-5} mol and 40.8 μmol each have three significant figures.

PRACTICE PROBLEMS

1. Express each of the following measurements in the units indicated.
   a. 2250 mg in grams
   b. 59.3 kL in liters

2. Use scientific notation to express each of the following measurements in the units indicated.
   a. 0.000 072 g in micrograms
   b. 3.98×10^{6} m in kilometers

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. A chemical compound always has the same elements in the same proportions by mass regardless of the source of the compound. This is a statement of
A. the law of multiple proportions.
B. the law of isotopes.
C. the law of definite proportions.
D. the law of conservation of mass.

2. An important result of Rutherford's experiments with gold foil was to establish that
A. atoms have mass.
B. electrons have a negative charge.
C. neutrons are uncharged particles.
D. the atom is mostly empty space.

3. Which subatomic particle has a charge of +1?
A. electron
B. neutron
C. proton
D. meson

4. Which particle has the least mass?
A. electron
B. neutron
C. proton
D. All have the same mass.

5. Cathode rays are composed of
A. alpha particles.
B. electrons.
C. protons.
D. neutrons.

6. The atomic number of an element is the same as the number of
A. protons.
B. neutrons.
C. protons + electrons.
D. protons + neutrons.

7. How many neutrons are present in an atom of tin that has an atomic number of 50 and a mass number of 119?
A. 50
B. 69
C. 119
D. 169

8. What is the mass of 1.50 mol of sodium, Na?
A. 0.652 g
B. 0.478 g
C. 11.0 g
D. 34.5 g

9. How many moles of carbon are in a 28.0 g sample?
A. 336 mol
B. 72.0 mol
C. 2.33 mol
D. 0.500 mol

SHORT ANSWER

10. Which atom has more neutrons, potassium-40 or argon-40?

11. What is the mass of 1.20×10^{23} atoms of phosphorus?

EXTENDED RESPONSE

12. Cathode rays emitted by a piece of silver and a piece of copper illustrate identical properties. What is the significance of this observation?

13. A student believed that she had discovered a new element and named it mythium. Analysis found it contained two isotopes. The composition of the isotopes was 19.9% of atomic mass 10.013 and 80.1% of atomic mass 11.009. What is the average atomic mass, and do you think mythium was a new element?

Test TIP Choose the best possible answer for each question, even if you think there is another possible answer that is not given.

Conservation of Mass

OBJECTIVES

- ***Observe*** **the signs of a chemical reaction.**
- ***Compare*** **masses of reactants and products.**
- ***Design*** **experiments.**
- ***Relate*** **observations to the law of conservation of mass.**

MATERIALS

- **2 L plastic soda bottle**
- **5% acetic acid solution (vinegar)**
- **balance**
- **clear plastic cups, 2**
- **graduated cylinder**
- **hook-insert cap for bottle**
- **microplunger**
- **sodium hydrogen carbonate (baking soda)**

FIGURE A Slowly add the vinegar to prevent the reaction from getting out of control.

BACKGROUND

The law of conservation of mass states that matter is neither created nor destroyed during a chemical reaction. Therefore, the mass of a system should remain constant during any chemical process. In this experiment, you will determine whether mass is conserved by examining a simple chemical reaction and comparing the mass of the system before the reaction with its mass after the reaction.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Make two data tables in your lab notebook, one for Part I and another for Part II. In each table, create three columns labeled “Initial mass (g),” “Final mass (g),” and “Change in mass (g).” Each table should also have space for observations of the reaction.

PROCEDURE—PART I

1. Obtain a microplunger, and tap it down into a sample of baking soda until the bulb end is packed with a plug of the powder (4–5 mL of baking soda should be enough to pack the bulb).
2. Hold the microplunger over a plastic cup, and squeeze the sides of the microplunger to loosen the plug of baking soda so that it falls into the cup.
3. Use a graduated cylinder to measure 100 mL of vinegar, and pour it into a second plastic cup.
4. Place the two cups side by side on the balance pan, and measure the total mass of the system

(before reaction) to the nearest 0.01 g. Record the mass in your data table.

5. Add the vinegar to the baking soda a little at a time to prevent the reaction from getting out of control, as shown in **Figure A.** Allow the vinegar to slowly run down the inside of the cup. Observe and record your observations about the reaction.

6. When the reaction is complete, place both cups on the balance, and determine the total final mass of the system to the nearest 0.01 g. Calculate any change in mass. Record both the final mass and any change in mass in your data table.

7. Examine the plastic bottle and the hook-insert cap. Try to develop a modified procedure that will test the law of conservation of mass more accurately than the procedure in Part I.

8. In your notebook, write the answers to items 1 through 3 in Analysis and Interpretation—Part I.

PROCEDURE—PART II

9. Your teacher should approve the procedure you designed in Procedure—Part I, step 7. Implement your procedure with the same chemicals and quantities you used in Part I, but use the bottle and hook-insert cap in place of the two cups. Record your data in your data table.

10. If you were successful in step 9 and your results reflect the conservation of mass, proceed to complete the experiment. If not, find a lab group that was successful, and discuss with them what they did and why they did it. Your group should then test the other group's procedure to determine whether their results are reproducible.

CLEANUP AND DISPOSAL

11. Clean your lab station. Clean all equipment, and return it to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or throw anything in the trash unless your teacher directs you to do so. Wash your hands thoroughly after all work is finished and before you leave the lab.

ANALYSIS AND INTERPRETATION—PART I

1. **Drawing Conclusions:** What evidence was there that a chemical reaction occurred?

2. **Organizing Data:** How did the final mass of the system compare with the initial mass of the system?

3. **Resolving Discrepancies:** Does your answer to the previous question show that the law of conservation of mass was violated? (Hint: Another way to express the law of conservation of mass is to say that the mass of all of the products equals the mass of all of the reactants.) What do you think might cause the mass difference?

ANALYSIS AND INTERPRETATION—PART II

1. **Drawing Conclusions:** Was there any new evidence in Part II indicating that a chemical reaction occurred?

2. **Organizing Ideas:** Identify the state of matter for each reactant in Part II. Identify the state of matter for each product.

CONCLUSIONS

1. **Relating Ideas:** What is the difference between the system in Part I and the system in Part II? What change led to the improved results in Part II?

2. **Evaluating Methods:** Why did the procedure for Part II work better than the procedure for Part I?

EXTENSIONS

1. **Applying Models:** When a log burns, the resulting ash obviously has less mass than the unburned log did. Explain whether this loss of mass violates the law of conservation of mass.

2. **Designing Experiments:** Design a procedure that would test the law of conservation of mass for the burning log described in Extension item 1.